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ABSTRACT

Teaching procedures of Project Physics Unit 5 are presented to help teachers make effective use of learning materials. Unit contents are discussed in connection with teaching aid lists, multi-media schedules, schedule blocks, and resource charts. Brief summaries are made for transparencies, 16mm films, and reader articles. Included is information about the background and development of each unit chapter, procedures used in demonstrations, apparatus operations, notes on the student handbook, and an explanation of film loops. Additional articles are concerned with relative atomic mass determination, spectroscopic experimentation, Rutherford scattering, angular momentum, and Nagaoka's theory of the "Saturnian" atom. A phototube unit and a Millikan setup are analyzed, and a bibliography of reference texts and periodicals is given. Solutions to the study guide are provided in detail, and answers to test items are suggested. The fifth unit of the text, with marginal notes on each section, is alsc compiled in the manual. The work of Harvard Project Physics has been financially supported by: the Carnegie Corporation of New York, the Ford Foundation, the National Science Foundation, the Alfred P. Sloan Foundation, the United States Office of Education, and Harvard University. (CC)

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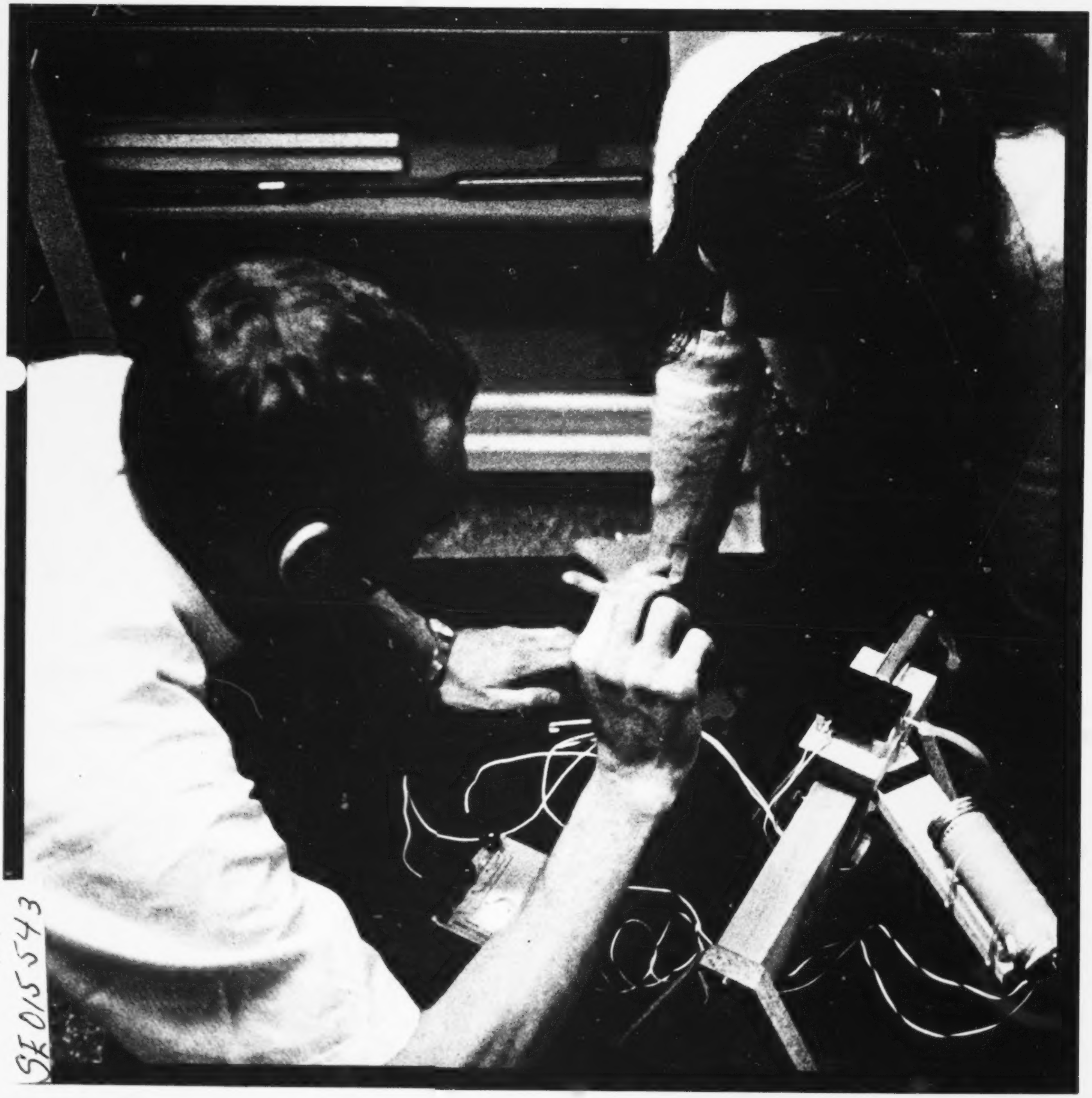
Project Physics **Teacher Guide** 5

An Introduction to Physics

Models of the Atom



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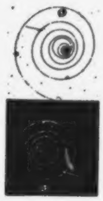
Project Physics **Text**

An Introduction to Physics 5 Models of the Atom



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Welcome to the study of physics. This volume, more of a student's guide than a text of the usual kind, is part of a whole group of materials that includes a student handbook, laboratory equipment, films, programmed instruction, readers, transparencies, and so forth. Harvard Project Physics has designed the materials to work together. They have all been tested in classes that supplied results to the Project for use in revisions of earlier versions.

The Project Physics course is the work of about 200 scientists, scholars, and teachers from all parts of the country, responding to a call by the National Science Foundation in 1963 to prepare a new introductory physics course for nationwide use. Harvard Project Physics was established in 1964, on the basis of a two-year feasibility study supported by the Carnegie Corporation. On the previous pages are the names of our colleagues who helped during the last six years in what became an extensive national curriculum development program. Some of them worked on a full-time basis for several years; others were part-time or occasional consultants, contributing to some aspect of the whole course; but all were valued and dedicated collaborators who richly earned the gratitude of everyone who cares about science and the improvement of science teaching.

Harvard Project Physics has received financial support from the Carnegie Corporation of New York, the Ford Foundation, the National Science Foundation, the Alfred P. Sloan Foundation, the United States Office of Education and Harvard University. In addition, the Project has had the essential support of several hundred participating schools throughout the United States and Canada, who used and tested the course as it went through several successive annual revisions.

The last and largest cycle of testing of all materials is now completed; the final version of the Project Physics course will be published in 1970 by Holt, Rinehart and Winston, Inc., and will incorporate the final revisions and improvements as necessary. To this end we invite our students and instructors to write to us if in practice they too discern ways of improving the course materials.

The Directors
Harvard Project Physics

An Introduction to Physics 5 Models of the Atom

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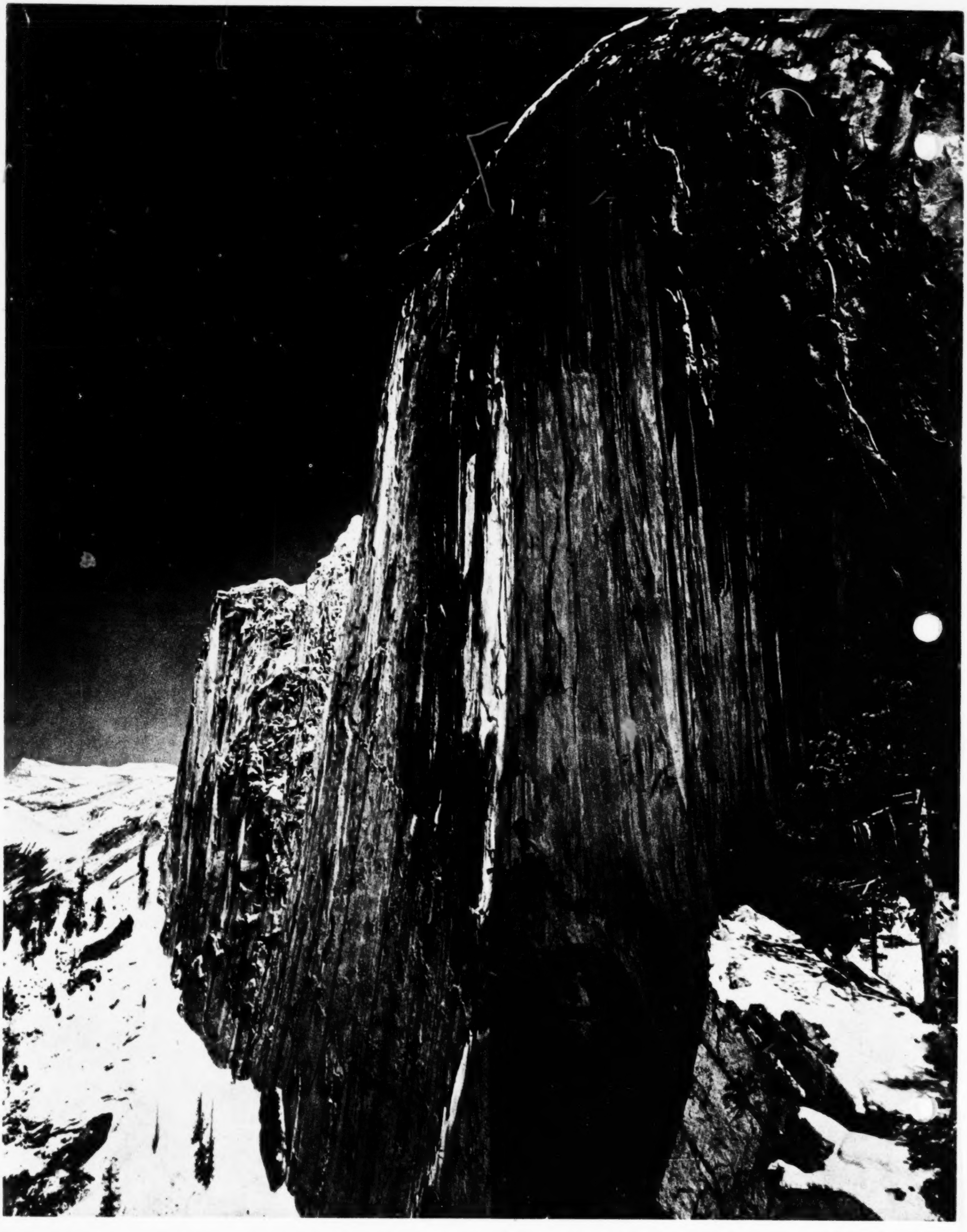


Prologue In the earlier units of this course we studied the motion of bodies: bodies of ordinary size, such as we deal with in everyday life, and very large bodies, such as planets. We have seen how the laws of motion and gravitation were developed over many centuries and how they are used. We have learned about conservation laws, about waves, about light, and about electric and magnetic fields. All that we have learned so far can be used to study a problem which has bothered people for many centuries: the problem of the nature of matter. The phrase, "the nature of matter," may seem simple to us now, but its meaning has been changing and growing over the centuries. The phrase really stands for the questions men ask about matter at any given date in the development of science; the kind of questions and the methods used to find answers to these questions are continually changing. For example, during the nineteenth century the study of the nature of matter consisted mainly of chemistry: in the twentieth century the study of matter has moved into atomic and nuclear physics.

Since 1800 progress has been so rapid that it is easy to forget that people have theorized about matter for more than 2500 years. In fact some of the questions for which answers have been found only during the last hundred years were asked more than two thousand years ago. Some of the ideas we consider new and exciting, such as the atomic constitution of matter, were debated in Greece in the fifth and fourth centuries B.C. In this prologue we shall, therefore, review briefly the development of ideas concerning the nature of matter up to about 1800. This review will set the stage for the four chapters of Unit 5, which will be devoted, in greater detail, to the progress made since 1800 on the problem of the constitution of matter. It will be shown in these chapters that matter is made up of atoms and that atoms have structures about which a great deal of information has been obtained.

Long before men started to develop the activities we call science, they were acquainted with snow, wind, rain, mist and cloud; with heat and cold; with salt and fresh water; wine, milk, blood and honey; ripe and unripe fruits; fertile and infertile seeds. They saw that plants, animals and men were born, that they grew and matured and that they aged and died. Men noticed that the world about them was continually changing and yet, on a large scale, it seemed to remain much the same. The causes of these changes and of the apparent continuity of nature were unknown. So men invented gods and demons who controlled nature. Myths grew up around the creation of the world and its contents, around

The photographs on these two pages illustrate some of the variety of forms of matter: large and small, stable and shifting, animate and inanimate.

microscopic crystals

condensed water vapor

toad on log

See "Structure, Substructure, Superstructure" in Project Physics Reader 5.

Monolith—The Face of Half Dome, 1927 (photo by Ansel Adams)

Summary

1. The first atomic theory of matter, introduced in Greece, was based upon a view of matter consisting of indivisible particles, made of the same substance.

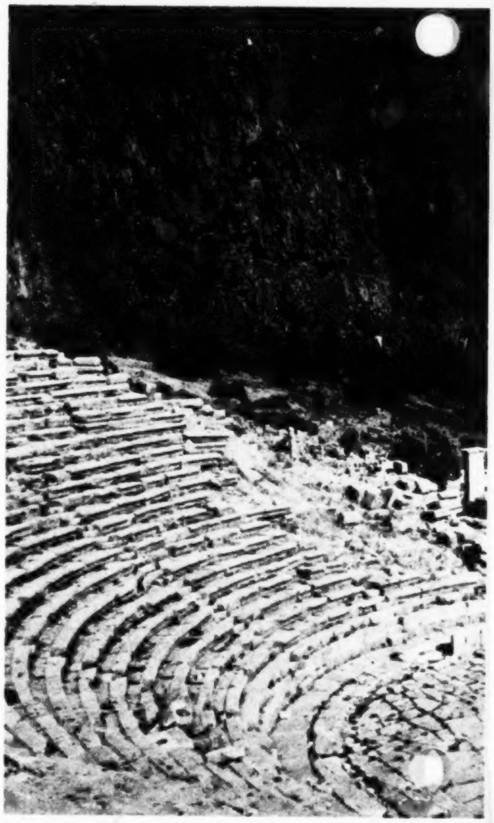
2. The atomic theory was criticized by Aristotle, who developed a theory based upon four elements: earth, air, fire and water.

3. Alchemy advanced the study of properties of substance, but left unsolved most of the fundamental problems.

4. The study of gases as well as increasingly precise observation of chemical reactions brought about a revival of atomism, and laid the groundwork for Dalton's atomic theory.

Greek Ideas of Order

The Greek mind loved clarity. In philosophy, literature, art and architecture it sought to interpret things with precision and in terms of their lasting qualities. It tried to discover the forms and patterns thought to be essential to an understanding of things. The Greeks delighted in showing these forms and patterns when they found them. Their art and architecture express beauty and intelligibility by means of clarity and balance of form. These aspects of Greek thought are beautifully expressed in the shrine of Delphi. The theater, which could seat 5,000 spectators, impresses us because of the size and depth of the tiered, semicircular seating structure. But even more striking is the balanced, orderly way in which the theater is shaped into the landscape so that the entire landscape takes on the aspect of a giant theater. The Athenian Treasury has an orderly system of proportions, with form and function integrated into a logical, pleasing whole. The statue of the charioteer, with its balance and firmness, represents a genuine ideal of male beauty. After more than 2,000 years we are still struck by the freedom and elegance of ancient Greek thought and expression.



Greek philosophers were the first to make a concerted effort to account for the universe in terms of natural law rather than divine intervention. Greek order was evident also in the concept of atoms. It is essentially Greek ideas which the Roman Lucretius expresses below.

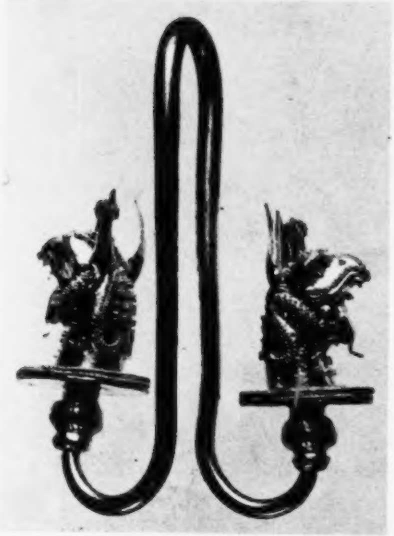
Basing his ideas on the tradition of atomists dating back to the greek philosophers, Democritus and Leucippus, Lucretius wrote in his poem, De rerum natura (Concerning the Nature of Things), "...Since the atoms are moving freely through the void, they must all be kept in motion either by their own weight or on occasion by the impact of another atom. For it must often happen that two of them in their course knock together and immediately bounce apart in opposite directions, a natural consequence of their hardness and solidity and the absence of anything behind to stop them.

"As a further indication that all particles of matter are on the move, remember that the universe is bottomless: there is no place where the atoms could come to rest. As I have already shown by various arguments and proved conclusively, space is without end or limit and spreads out immeasurably in all directions alike.

"It clearly follows that no rest is given to atoms in their course through the depths of space. Driven along in an incessant but variable movement, some of them bounce far apart after a collision while others recoil only a short distance from the impact. From those that do not recoil far, being driven into a closer union and held there by the entanglement of their own interlocking shapes, are composed firmly rooted rock, the stubborn strength of steel and the like. Those others that move freely through larger tracts of space, springing far apart and carried far by the rebound—these provide for us thin air and blazing sunlight. Besides these, there are many other atoms at large in empty space which have been thrown out of compound bodies and have nowhere even been granted admittance so as to bring their motions into harmony."

the changes of the seasons, around the events man could see happening but could not understand.

Over a long period of time men developed some control over nature: they learned how to smelt ores, to make weapons and tools, to produce gold ornaments, glass, perfumes, medicines and beer. Eventually, in Greece, about the year 600 B.C., philosophers—the lovers of wisdom—started to look for rational explanations of natural events, that is, explanations that did not depend on the whims of gods or demons. They sought to discover the enduring, unchanging things out of which the world is made. They started on the problem of explaining how these enduring things can give rise to the changes that we perceive. This was the beginning of man's attempts to understand the material world—the nature of matter.



This gold earring, made in Greece about 600 B.C., shows the great skill with which ancient artisans worked metals.
Museum of Fine Arts, Boston.

The earliest Greek philosophers thought that all the different things in the world were made out of a single basic substance, or stuff. Some thought that water was the fundamental substance and that all other substances were derived from it. Others thought that air was the basic substance; still others favored fire. But neither water, air nor fire was satisfactory; no one substance seemed to have enough different properties to give rise to the enormous variety of substances in the world. According to another view, introduced by Empedocles around 450 B.C., there are four basic types of matter: earth, air, fire and water; all material things were made out of them. Change comes about through the mingling and separation of these four basic materials which unite in different proportions to produce the familiar objects around us; but the basic materials were supposed to persist through all these changes. This theory was the first appearance in our scientific tradition of a model of matter according to which all material things are just different arrangements of a few eternal substances, or elements.

Leucippus arrived at the concept of the atom by rejecting the philosophical notion that matter was infinitely divisible. He considered it inconceivable that matter could be divided without limit into ever smaller pieces. There must be a smallest piece which cannot be divided further.

The first atomic theory of matter was introduced by the Greek philosopher Leucippus, born about 500 B.C., and his pupil Democritus, who lived from about 460 B.C. to 370 B.C. Only scattered fragments of the writings of these philosophers remain, but their ideas are discussed in considerable detail by Aristotle (384-322 B.C.), by another Greek philosopher, Epicurus (341-270 B.C.) and by the Latin poet Lucretius (100-55 B.C.). It is to these men that we owe most of our knowledge of ancient atomism.

The theory of the atomists was based on a number of assumptions: (1) that matter is eternal, and that no material thing can come from nothing, nor can anything material pass into nothing; (2) that material things consist of very minute, but not infinitely small, indivisible particles—the word "atom" meant "uncuttable" in Greek and, in discussing the ideas of the early atomists, we could use the word "indivisibles" instead of the word "atoms"; (3) that all atoms are of the same kind, that is, of the same substance, but differ in size, shape and position; (4) that the atoms exist in otherwise empty space (void), which separates them, and because of this space they are capable of movement; (5) that the atoms are in ceaseless motion although the nature and cause of the atomic motions are not clear. In the course of their motions atoms come together and form combinations which are the material substances we know. When the atoms forming these combinations separate, the substances break up. Thus, the combinations and separations of atoms give rise to the changes which take place in the world. The combinations and separations take place in accord with natural laws which are not known, but do not require the action of gods or demons or other supernatural powers.

With the above assumptions, the ancient atomists were able to work out a consistent story of change, of what they sometimes called "coming-to-be" and "passing-away." They could not prove experimentally that their theory was correct, and they had to be satisfied with a rational explanation based on assumptions that seemed reasonable to them. The theory was a "likely story," but it was not useful for the prediction of new phenomena.

The atomic theory was criticized severely by Aristotle, who argued, on logical grounds, that no vacuum or void could exist and that the ideas of atoms with their inherent motion must be rejected. For a long time Aristotle's argument against the void was convincing. Not until the seventeenth century did Torricelli's experiments (described in Chapter 12) show that a vacuum could indeed exist. Aristotle also argued that matter is continuous and infinitely divisible so that there can be no atoms.

Aristotle developed a theory of matter as part of his grand scheme of the universe, and this theory, with some modifications, was thought to be satisfactory by most philosophers of nature for nearly two thousand years. His theory of matter was based on the four basic substances

Impressive as Greek insight may seem, no real evidence in support of an atomic theory was produced until the end of the eighteenth century.

According to Aristotle in his Metaphysics, "There is no consensus concerning the number or nature of these fundamental substances. Thales, the first to think about such matters, held that the elementary substance is clear liquid....He may have gotten this idea from the observation that only moist matter can be wholly integrated into an object—so that all growth depends on moisture....

"Anaximenes and Diogenes held that colorless gas is more elementary than clear liquid, and that, indeed, it is the most elementary of all simple substances. On the other hand, Hippasus of Metapontum and Heraclitus of Ephesus said that the most elementary substance is heat. Empedocles spoke of four elementary substances, adding dry dust to the three already mentioned...Anaxagoras of Clazomenae says that there are an infinite number of elementary constituents of matter...." [From a translation by D. E. Gershenson and D. A. Greenberg.]

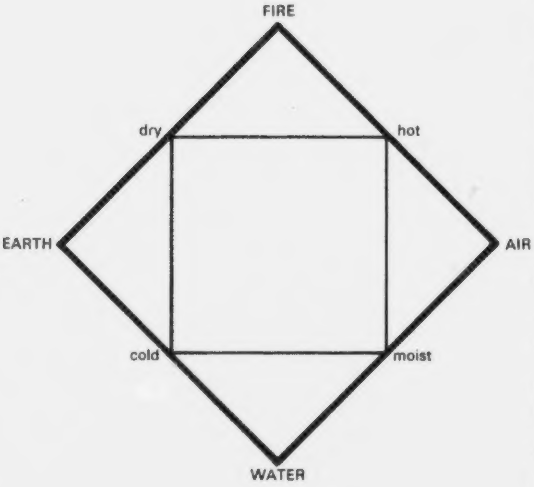
It should be emphasized that Aristotle's theory of matter met many of the demands of a good theory.

or "elements," earth, air, fire and water, and four "qualities," cold, hot, moist and dry. Each element was characterized by two qualities. Thus the element

earth is dry and cold
water is cold and moist
air is moist and hot
fire is hot and dry.

According to Aristotle, it is always the first of the two qualities which predominates. The elements are not unchangeable; any one of them may be transformed into any other because of one or both of its qualities changing into opposites. The transformation takes place most easily between two elements having one quality in common; thus earth is transformed into water when dryness changes into moistness. Aristotle worked out a scheme of such possible transformations which can be shown in the following diagram:

The four elements were not the same as ordinary earth, air, fire and water, each of which had some admixture of all.



Earth can also be transformed into air if both of the qualities of earth (dry, cold) are changed into their opposites (moist, hot). Water can be transformed into fire if both of its qualities (cold, moist) are changed into their opposites (hot, dry).

Aristotle was also able to explain many natural phenomena by means of his ideas. Like the atomic theory, Aristotle's theory of coming-to-be and passing-away was consistent, and constituted a model of the nature of matter. It had certain advantages over the atomic theory: it was based on elements and qualities that were familiar to people; it did not involve the use of atoms, which couldn't be seen or otherwise perceived, or of a void, which was difficult to imagine. In addition, Aristotle's theory provided some basis for further experimentation: it supplied what seemed like a rational basis for the possibility of changing one material into another.

During the period 300 A.D. to about 1600 A.D., atomism declined although it did not die out completely. Christian, Hebrew and Moslem theologians considered atomists to be "atheistic" and "materialistic" because they claimed that everything in the universe can be explained in terms of matter and motion. The atoms of Leucippus and Democritus moved through empty space, devoid of spirit, and with no definite plan or purpose. Such an idea was contrary to the beliefs of the major religions.

About 300 or 400 years after Aristotle, a kind of research called alchemy appeared in the Near and Far East. Alchemy in the Near East was a combination of Aristotle's ideas about matter with methods of treating ores and metals. One of the aims of the alchemists was to change, or transmute, ordinary metals into gold. Although they failed to do this, alchemy (along with metallurgy) was a forerunner of chemistry. The alchemists studied many of the properties of substances that are now classified as chemical properties. They invented many of the pieces of chemical apparatus that are still used, such as reaction vessels (retorts) and distillation flasks. They studied such processes as calcination, distillation, fermentation and sublimation. In this sense alchemy may be regarded as the chemistry of the Middle Ages. But alchemy left unsolved some of the fundamental questions. At the opening of the eighteenth century the most important of these questions were: first, what is a chemical element; second, what is the nature of chemical composition and chemical change, especially burning; third, what is the chemical nature of the so-called elements, air, fire and water. Until these questions were answered, it was impossible to make real progress in finding out what matter is. One result was that the "scientific revolution" of the seventeenth century, which clarified the problems of astronomy and dynamics, did not reach chemistry until the eighteenth century.

Laboratory of a 16th-century alchemist.

Hydrogen and oxygen were unknown to mediaeval alchemists.

During the seventeenth century, however, some forward steps were made which supplied a basis for future progress on the problem of matter. The Copernican and Newtonian revolutions undermined the authority of Aristotle to such an extent that his ideas about matter were also questioned. Atomic concepts were revived because atomism offered a way of looking at things that was very different from Aristotle's ideas. As a result theories involving "atoms," "particles" or "corpuscles" were again considered seriously. Boyle's models of a gas (Chapter 11) were based on the idea of "gas particles." Newton also discussed the behavior of a gas (and even of light!) by supposing it to consist of particles. Thus, the stage was set for a general revival of atomic theory.

One of those who contributed greatly to the revival of atomism was Pierre Gassendi (1592-1655), a French priest and philosopher. He avoided the criticism of atomism as atheistic by saying that God also created the atoms and bestowed motion upon them. Gassendi accepted the physical explanations of the atomists, but rejected their disbelief in the immortality of the soul and in Divine Providence. He was thus able to provide a philosophical justification of atomism which met some of the serious religious objections.

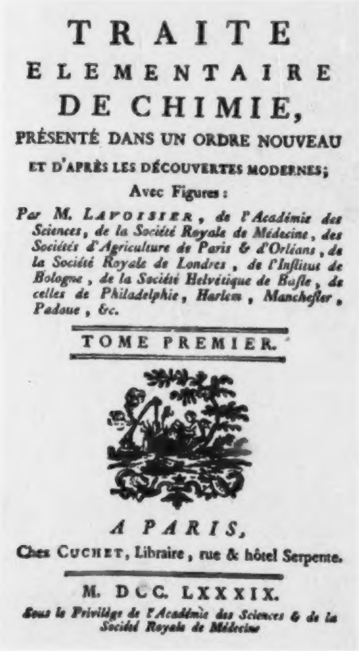
In the eighteenth century, chemistry became more quantitative as the use of the balance was increased. Many new substances were isolated and their properties examined. The attitude that grew up in the latter half of the century was exemplified by that of Henry Cavendish (1731-1810), who, according to a biographer, regarded the universe as consisting

> ...solely of a multitude of objects which could be weighed, numbered, and measured; and the vocation to which he considered himself called was to weigh, number, and measure as many of those objects as his alloted threescore years and ten would permit....He weighed the Earth; he analysed the Air; he discovered the compound nature of Water; he noted with numerical precision the obscure actions of the ancient element Fire.

A Swedish pharmacist, Carl W. Scheele, discovered oxygen before Priestley. But his results did not appear in print until after Priestley's independent work was published.

Eighteenth-century chemistry reached its peak in the work of Lavoisier (1743-1794), who worked out the modern views of combustion, established the law of conservation of mass (see Chapter 9), explained the elementary nature of hydrogen and oxygen and the composition of water, and emphasized the quantitative aspects of chemistry. His famous book, Traité Elémentaire de Chimie (or Elements of Chemistry), published in 1789, established chemistry as a modern science. In it, he analyzed the idea of element in a way which is very close to our modern views:

> ...if, by the term elements we mean to express those simple and indivisible atoms of which matter is composed, it is extremely probable that we know nothing at all about them; but if we apply the term elements, or principles of bodies, to express our idea of the last point which analysis is capable of reaching, we must admit as elements, all the substances into which we are capable, by any means, to reduce bodies by decomposition. Not that we are entitled to affirm that these substances we consider as simple may not be compounded of two, or even of a greater number of principles; but since these principles cannot be separated, or rather since we have not hitherto discovered the means of separating them, they act with regard to us as simple substances, and we ought never to suppose them compounded until experiment and observation have proved them to be so.



TRAITE
ELEMENTAIRE
DE CHIMIE,
PRÉSENTÉ DANS UN ORDRE NOUVEAU
ET D'APRÈS LES DÉCOUVERTES MODERNES;
Avec Figures:
Par M. Lavoisier, de l'Académie des Sciences, de la Société Royale de Médecine, des Sociétés d'Agriculture de Paris & d'Orléans, de la Société Royale de Londres, de l'Inſtitut de Bologne, de la Société Helvétique de Baſle, de celles de Philadelphie, Harlem, Mancheſter, Padoue, &c.

TOME PREMIER.

A PARIS,
Chez Cuchet, Libraire, rue & hôtel Serpente.

M. DCC. LXXXIX.
Sous le Privilège de l'Académie des Sciences & de la Société Royale de Médecine

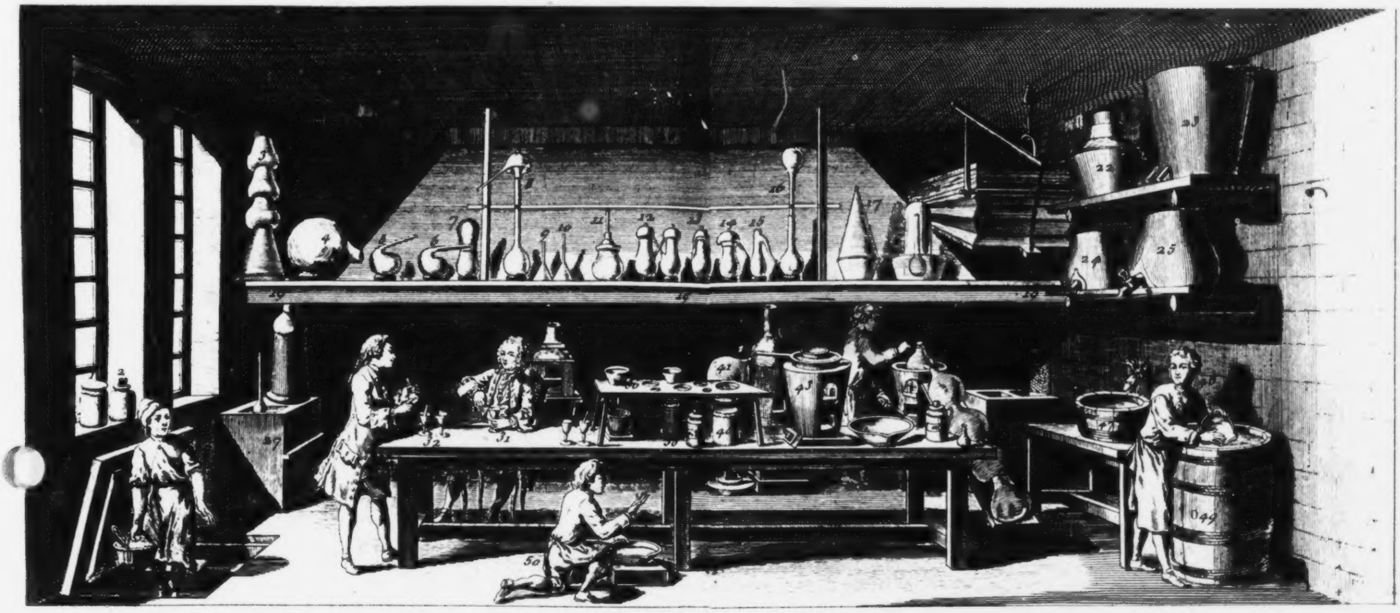
Title page of Lavoisier's Traité Elémentarie de Chimie (1789)

During the latter half of the eighteenth century and the early years of the nineteenth century great progress was made in chemistry because of the increasing use of quantitative methods. Chemists found out more and more about the composition of substances. They separated many elements and showed that nearly all substances are compounds—combinations—of chemical elements. They learned a great deal about how elements combine to form compounds and how compounds can be broken down into the elements of which they are composed. This information made it possible for chemists to establish certain laws of chemical combination. Then chemists sought an explanation for these laws.

During the first ten years of the nineteenth century, John Dalton, an English chemist, introduced a modified form of the old Greek atomic theory to account for the laws of chemical combination. It is here that the modern story of the atom begins. Dalton's theory was an improvement over that of Democritus, Epicurus and Lucretius because it opened the way for the quantitative study of the atom in the nineteenth century. Today the existence of the atom is no longer a topic of speculation. There are many kinds of experimental evidence, not only for the existence of atoms but also for their structure. This evidence, which began to accumulate about 150 years ago, is now convincing. In this unit we shall trace the discoveries and ideas that produced this evidence

The first mass of convincing evidence for the existence of atoms and the first clues to the nature of atoms came from chemistry. We shall, therefore, start with chemistry in the early years of the nineteenth century; this is the subject of Chapter 17. We shall see that chemistry raised certain questions about atoms which could only be answered by physics. Physical evidence, accumulated in the nineteenth century and the early years of the twentieth century, made it possible to propose atomic models of atomic structure. This evidence and the earlier models will be discussed in Chapters 18 and 19. The latest ideas about atomic theory will be discussed in Chapter 20.

A chemical laboratory of the 18th century



Chapter 17 The Chemical Basis of Atomic Theory

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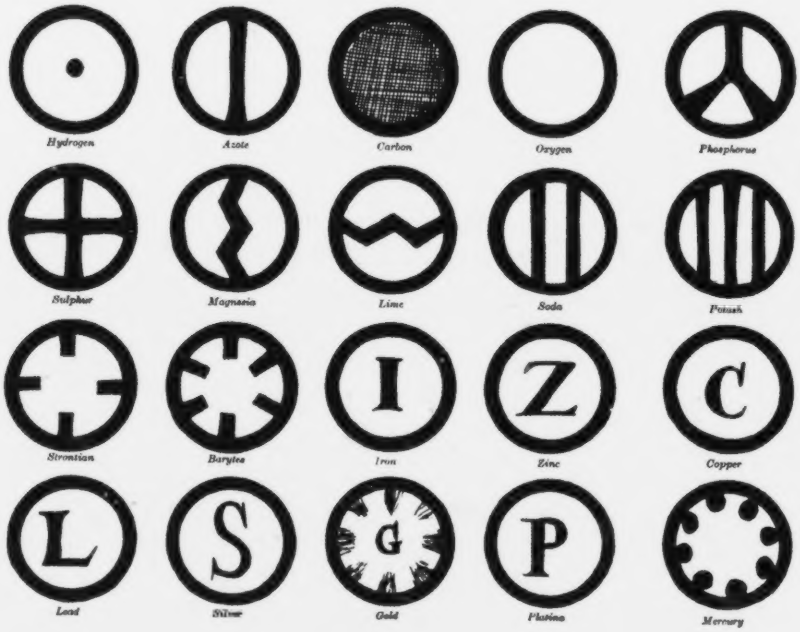


Hydrogen
Azote
Carbon
Oxygen
Phosphorus
Sulphur
Magnesia
Lime
Soda
Potash
Strontian
Barytes
I
Iron
Z
Zinc
C
Copper
L
Lead
S
Silver
G
Gold
P
Platina
Mercury

Dalton's symbols of the elements

17.1 Dalton's atomic theory and the laws of chemical combination.

The atomic theory of John Dalton appeared in his treatise, A New System of Chemical Philosophy, published in two parts, in 1808 and 1810. The main postulates of his theory were:

(1) Matter consists of indivisible atoms.

> ...matter, though divisible in an extreme degree, is nevertheless not infinitely divisible. That is, there must be some point beyond which we cannot go in the division of matter. The existence of these ultimate particles of matter can scarcely be doubted, though they are probably much too small ever to be exhibited by microscopic improvements. I have chosen the word atom to signify these ultimate particles....

(2) Each element consists of a characteristic kind of identical atoms. There are consequently as many different kinds of atoms as there are elements. The atoms of an element "are perfectly alike in weight and figure, etc."

(3) Atoms are unchangeable.

(4) When different elements combine to form a compound, the smallest portion of the compound consists of a grouping of a definite number of atoms of each element.

(5) In chemical reactions, atoms are neither created nor destroyed, but only rearranged.

Summary 17.1

1. The postulates of Dalton's atomic theory form the basis for modern chemistry. From his postulates one can deduce
a) the law of conservation of mass.
b) the law of definite proportions.
(Nonetheless, all but number 5 are now known to be not strictly true.)

Dalton's theory really grew out of his interest in meteorology and his research on the composition of the atmosphere. He tried to explain many of the physical properties of gases in terms of atoms. At first he assumed that the atoms of all the different elements had the same size. But this assumption didn't work and he was led to think of the atoms of different elements as being different in size or in mass. In keeping with the quantitative spirit of the time, he tried to determine the numerical values for the differences in mass. But before considering how to determine the masses of atoms of the different elements, let us see how Dalton's postulates make it possible to account for the laws of chemical combination.

Meteorology is a science that deals with the atmosphere and its phenomena—weather forecasting is one branch of meteorology.

Dalton's theory was pure speculation, as was Democritus'. But Dalton's was based on an extensive body of a quantitative knowledge of chemical reactions and it was more detailed and explicit in describing combinations of atoms.

We consider first the law of conservation of mass. In 1774, Lavoisier studied the reaction between tin and oxygen in a closed and sealed container. When the tin is heated in air, it reacts with the oxygen in the air to form a compound, tin oxide, which is a white powder. Lavoisier weighed the sealed container before and after the chemical reaction and found that the mass of the container and its contents was the same before and after the reaction. A modern example of a similar reaction is the flashing of a photographic flash bulb containing magnesium. The flash bulb is an isolated system containing two elements, magnesium (in the form of a wire) and oxygen gas, sealed in a closed container. When an electric current passes through the wire, a chemical reaction occurs with a brilliant flash. Magnesium and oxygen disappear, and a white powder, magnesium oxide, is formed. Comparison of the mass after the reaction with the mass before the reaction shows that there is no detectable change in mass; the mass is the same before and after the reaction. Careful work by many experimenters on many chemical reactions has shown that mass is neither destroyed nor created, in any detectable amount, in a chemical reaction. This is the law of conservation of mass.

F35 Definite and multiple proportions.

F36 Elements, compounds and mixtures

See "Failure and Success" in Project Physics Reader 5.

According to Dalton's theory (postulates 4 and 5) chemical changes are only the rearrangements of unions of atoms. Since atoms are unchangeable (according to postulate 3) rearranging them cannot change their masses. Hence, the total mass of all the atoms before the reaction must equal the total mass of all the atoms after the reaction. Dalton's atomic theory, therefore, accounts in a simple and direct way for the law of conservation of mass.

A second law of chemical combination which could be explained easily with Dalton's theory is the law of definite proportions. This law states that a particular chemical compound always contains the same elements united in the same proportions by weight. For example, the ratio of the masses of oxygen and hydrogen which combine to form water is always 7.94 to 1, that is,

$$\frac{\text{mass of oxygen}}{\text{mass of hydrogen}} = \frac{7.94}{1}.$$

If there is more of one element present, say 10 grams of oxygen and one gram of hydrogen, only 7.94 grams of oxygen will combine with the hydrogen. The rest of the oxygen, 2.06 grams, remains uncombined.

The fact that elements combine in fixed proportions means also that each chemical compound has a certain definite composition. Thus, by weight, water consists of 88.8 percent oxygen and 11.2 percent hydrogen. The decomposition of sodium chloride (common salt) always gives the results: 39 percent sodium and 61 percent chlorine by weight. This is another way of saying that 10 grams of sodium always combine with 15.4 grams of chlorine to form sodium chloride. Hence, the law of definite proportions is also referred to as the law of definite composition.

15.6 grams of chlorine would be better.

Now let us see how Dalton's theory can be applied to *a chemical reaction, say, to the formation of water from* oxygen and hydrogen. First, according to Dalton's second postulate, all the atoms of oxygen have the same mass; all the atoms of hydrogen have the same mass, which is different from the mass of the oxygen atoms. To express the mass of the oxygen entering into the reaction, we multiply the mass of a single oxygen atom by the number of oxygen atoms:

$$\text{mass of oxygen} = \begin{pmatrix}\text{mass of 1}\\ \text{oxygen atom}\end{pmatrix} \times \begin{pmatrix}\text{number of}\\ \text{oxygen atoms}\end{pmatrix}.$$

SG 17.1

SG 17.2

Similarly, the mass of hydrogen entering into the reaction is equal to the product of the number of hydrogen atoms entering into the reaction and the mass of one hydrogen atom:

$$\text{mass of hydrogen} = \begin{pmatrix}\text{mass of 1}\\ \text{hydrogen atom}\end{pmatrix} \times \begin{pmatrix}\text{number of}\\ \text{hydrogen atoms}\end{pmatrix}.$$

Symbols for atoms of hydrogen and oxygen were introduced in 1813 by the Swedish chemist Berzelius.

To get the ratio of the mass of oxygen entering into the reaction to the mass of hydrogen entering into the reaction, we divide the first equation by the second equation:

$$\frac{\text{mass of oxygen}}{\text{mass of hydrogen}} = \frac{\begin{pmatrix}\text{mass of 1}\\ \text{oxygen atom}\end{pmatrix}}{\begin{pmatrix}\text{mass of 1}\\ \text{hydrogen atom}\end{pmatrix}} \times \frac{\begin{pmatrix}\text{number of}\\ \text{oxygen atoms}\end{pmatrix}}{\begin{pmatrix}\text{number of}\\ \text{hydrogen atoms}\end{pmatrix}}.$$

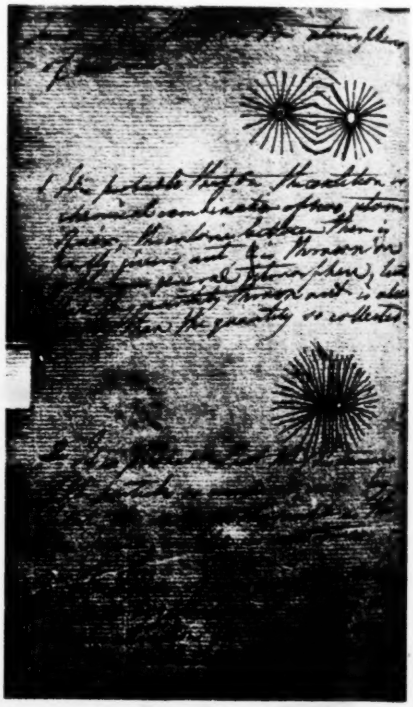
SG 17.3

Now, the masses of the atoms do not change (postulate 3), so the first ratio on the right side of the resulting equation has a certain unchangeable value. According to postulate 4, the smallest portion of the compound, water (now called a molecule of water) consists of a definite number of atoms of each element. Hence the numerator of the second ratio on the right side of the equation has a definite value, and the denominator has a definite value, so the ratio has a definite value. The product of the two ratios on the right hand side therefore, has a certain definite value. This equation then tells us that the ratio of the masses of oxygen and hydrogen that combine to form water must have a certain definite value. But this is just the law of definite proportions or definite composition. Thus, Dalton's theory also accounts for this law of chemical combination.

See TG, p. 87 for comments on the determination of relative atomic masses.

There are other laws of chemical combination which are explained by Dalton's theory. Because the argument would

John Dalton (1766-1844). His first love was meteorology and he kept careful daily weather records for 46 years—a total of 200,000 observations. He was the first to describe color blindness in a publication, and was color-blind himself, not exactly an advantage for a chemist who had to see color changes in chemicals (his color blindness may help to explain why Dalton was a rather clumsy and slipshod experimenter). But his accomplishments rest not upon successful experiments, but upon his interpretation of the work of others. Dalton's notion that all elements were composed of extremely tiny, indivisible and indestructible atoms, and that all substances are composed of combinations of these atoms was accepted by most chemists with surprisingly little opposition. There were many attempts to honor him, but being a Quaker, he shunned any form of glory. When he received a doctor's degree from Oxford, his colleagues wanted to present him to King William IV. He had always resisted such a presentation because he would not wear court dress. However, his Oxford robes would satisfy the protocol. Unfortunately, they were scarlet and a Quaker could not wear scarlet. But Dalton could see no scarlet and was presented to the king in robes which he saw as gray.



A page from Dalton's notebook, showing his representation of two adjacent atoms (top) and of a molecule or "compound atom" (bottom)

get complicated and nothing really new would be added to our argument, we shall not discuss them.

Dalton's interpretation of the experimental facts of chemical combination made possible several important conclusions: (1) that the differences between one chemical element and another would have to be described in terms of the differences between the atoms of which these elements were made up; (2) that there were, therefore, as many different types of atoms as there were chemical elements; (3) that chemical combination was the union of atoms of different elements into molecules of compounds. Dalton's theory also showed that the analysis of a large number of chemical compounds could make it possible to assign relative mass values to the atoms of different elements. This possibility will be discussed in the next section.

Q1 What did Dalton assume about the atoms of an element?

Q2 What two experimental laws did Dalton's assumptions explain?

17.2 The atomic masses of the elements. One of the most important concepts to come from Dalton's work is that of atomic mass and the possibility of determining numerical values for the masses of the atoms of different elements. Dalton had no idea of the actual masses of atoms except that he thought they were very small. In addition, reasonable estimates of atomic size did not appear until about 50 years after Dalton published his theory. They came from the kinetic theory of gases and indicated that atoms (or molecules) had diameters of the order of 10^{-10} meter. Atoms are thus much too small for mass measurements to be made on single atoms. But relative values of atomic masses can be found by using the law of definite proportions and experimental data on chemical reactions.

To see how this could be done we return to the case of water, for which, as we saw in the last section, the ratio of the mass of oxygen to the mass of hydrogen is 7.94:1. Now, if we knew how many atoms of oxygen and hydrogen are contained in a molecule of water we could find the ratio of the mass of the oxygen atom to the mass of the hydrogen atom. Dalton didn't know the numbers of oxygen and hydrogen atoms in a molecule of water. He therefore made an assumption. As scientists often do, he made the simplest assumption, namely, that one atom of oxygen combines with one atom of hydrogen to form one "compound atom" (molecule) of water. By this reasoning Dalton concluded that the oxygen atom is 7.94 times more massive than the hydrogen atom.

Summary 17.2
1. Although the actual atomic masses could not be determined by using only the law of definite proportions plus some simple assumptions, the relative atomic masses could.

2. Dalton first chose the atomic mass of hydrogen as the unit for the scale of atomic masses. Later Oxygen was used (as 16) instead of hydrogen for this purpose, and now carbon 12 is used.

SG 17.4
SG 17.5

More generally, Dalton assumed that when only one compound of two elements, A and B, exists, one atom of A always combines with one atom of B. Although Dalton could then find values of the relative masses of different atoms later work showed that Dalton's assumption of one to one ratios was often incorrect. For example, it was found that one atom of oxygen combines with two atoms of hydrogen to form one molecule of water, so the ratio of the mass of an oxygen atom to the mass of a hydrogen atom is 15.88 instead of 7.94. By studying the composition of water as well as many other chemical compounds, Dalton found that the hydrogen atom appeared to have less mass than the atoms of any other elements. Therefore, he proposed to express the masses of atoms of all other elements in terms of the mass of the hydrogen atom. Dalton defined the relative atomic mass of an element as the mass of an atom of that element compared to the mass of a hydrogen atom. This definition could be used by chemists in the

Avogadro corrected the errors due to Dalton's rule of simplicity.

Although "atomic weight" is commonly used by chemists (who have done all their work close to the earth's surface), we have tried to use the more correct "atomic mass" everywhere save in historical quotations.

nineteenth century even before the masses of individual atoms could be measured directly. All that was needed was the ratios of masses of atoms; these ratios could be found by measuring the masses of substances in chemical reactions (see Sec. 17.1). For example, we can say that the mass of a hydrogen atom is "one atomic mass unit" (1 amu). Then, if we know that an oxygen atom has a mass 15.88 times as great as that of a hydrogen atom, we can say that the atomic mass of oxygen is 15.88 atomic mass units. The system of atomic masses used in modern physical science is based on this principle, although it differs in details (and the standard for comparison is now carbon• instead of oxygen).

SG 17.6
SG 17.7

•Because carbon forms compounds with a great variety of molecular weights, giving a good choice for practical comparison, especially in a mass spectrometer.

During the nineteenth century chemists extended and improved Dalton's ideas. They studied many chemical reactions quantitatively, and developed highly accurate methods for determining relative atomic and molecular masses. More elements were isolated and their relative atomic masses determined. Because oxygen combined readily with many other elements chemists decided to use oxygen rather than hydrogen as the standard for atomic masses. Oxygen was assigned an atomic mass of 16 so that hydrogen could have an atomic mass close to one. The atomic masses of other elements could be obtained, relative to that of oxygen, by applying the laws of chemical combination to the compounds of the elements with oxygen. By 1872, 63 elements had been identified and their atomic masses determined. They are listed in Table 17.1, which gives modern values for the atomic masses. This table contains much valuable information, which we shall consider at greater length in Sec. 17.4. (The special marks, circles and rectangles, will be useful then.)

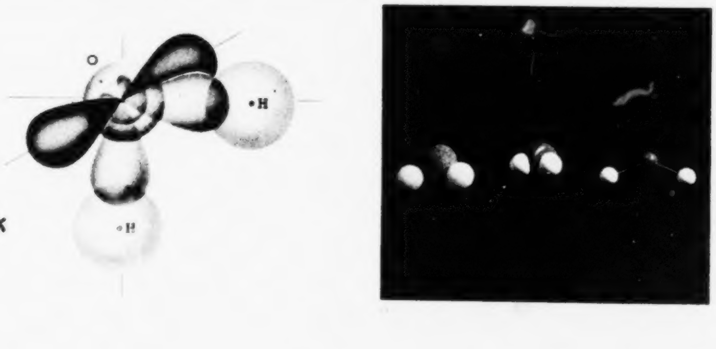
The progress made in identifying elements in the 19th century may be seen in the following table.

Year	Total number of elements identified
1720	14
1740	15
1760	17
1780	21
1800	31
1820	49
1840	56
1860	60
1880	69
1900	83

Q3 Was the simplest chemical formula necessarily correct?

Q4 Why did Dalton choose hydrogen as the unit of atomic mass?

Several different representations of a water molecule.



From CBA book

Table 17.1 Elements known by 1872

Name	Symbol	Atomic Mass*
hydrogen	H	1.0
□ lithium	Li	6.9
beryllium	Be	9.0
boron	B	10.8
carbon	C	12.0
nitrogen	N	14.0
oxygen	O	16.0
○ fluorine	F	19.0
□ sodium	Na	23.0
magnesium	Mg	24.3
aluminum	Al	27.0
silicon	Si	28.1
phosphorus	P	31.0
sulfur	S	32.1
○ chlorine	Cl	35.5
□ potassium	K	39.1
calcium	Ca	40.1
titanium	Ti	47.9
vanadium	V	50.9
chromium	Cr	52.0
manganese	Mn	54.9
iron	Fe	55.8
cobalt	Co	58.9
nickel	Ni	58.7
copper	Cu	63.5
zinc	Zn	65.4
arsenic	As	74.9
selenium	Se	79.0
○ bromine	Br	79.9
□ rubidium	Rb	85.5
strontium	Sr	87.6
yttrium	Yt	88.9
zirconium	Zr	91.2
niobium	Nb	92.9
molybdenum	Mo	95.9
ruthenium	Ru	101.1(104)
rhodium	Rh	102.9(104)
palladium	Pd	106.4
silver	Ag	107.9
cadmium	Cd	112.4
indium	In	114.8(113)
tin	Sn	118.7
antimony	Sb	121.7
tellurium	Te	127.6(125)
○ iodine	I	126.9
□ cesium	Cs	132.9
barium	Ba	137.3
didymium(**)	Di	—— (138)
cerium	Ce	140.1
erbium	Er	167.3(178)
lanthanum	La	138.9(180)
tantalum	Ta	180.9(182)
tungsten	W	183.9
osmium	Os	190.2(195)
iridium	Ir	192.2(197)
platinum	Pt	195.1(198)
gold	Au	197.0(199)
mercury	Hg	200.6
thallium	Tl	204.4
lead	Pb	207.2
bismuth	Bi	209.0
thorium	Th	232.0
uranium	U	238.0(240)

* Atomic masses given are modern values. Where these differ greatly from those accepted in 1872, the old values are given in parentheses.

** Didymium (Di) was later shown to be a mixture of two different elements, namely praseodymium (Pr; atomic mass 140.9) and neodymium (Nd; atomic mass 144.2).

○halogens □alkaline metals

17.3 Other properties of the elements: valence. In addition to the atomic masses, many other properties of the elements and their compounds were determined. Among these properties were: melting point, boiling point, density, electrical conductivity, thermal conductivity (the ability to conduct heat), specific heat (the amount of heat needed to change the temperature of one gram of a substance by 1°C), hardness, refractive index and others. The result was that by 1870 an enormous amount of information was available about a large number of elements and their compounds.

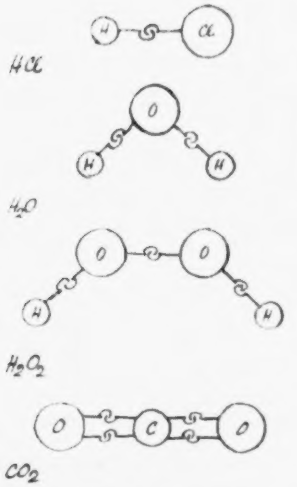
One of the most important properties that chemists studied was the combining ability or combining capacity of an element. This property, which is called valence,

In the thirteenth century a great theologian and philosopher, Albertus Magnus (Albert the Great) introduced the idea of affinity to denote an attractive force between substances that causes them to enter into chemical combination. It was not until 600 years later that it became possible to replace this qualitative notion by quantitative concepts. Valence is one of these quantitative concepts.

Summary 17.3

1. Each element can be characterized by its ability to combine with other elements.

2. The combining capacity of an element is known as its "valence" and is represented by a number. Some elements require more than one valence to characterize their combining capacities.



Representations of molecules formed from "atoms with hooks."

element	H	Cl	O	C
valence #	1	1	2	4

plays an important part in our story. As a result of studies of chemical compounds, chemists were able to assign formulas to the molecules of compounds. These formulas show how many atoms of each element are contained in a molecule. For example, water has the familiar formula H_2O, which indicates that the smallest piece of water that exists as water contains two atoms of hydrogen and one atom of oxygen. Hydrogen chloride (hydrochloric acid) has the formula HCl; one atom of hydrogen combines with one atom of chlorine. Common salt may be represented by the formula NaCl; this indicates that one atom of sodium combines with one atom of chlorine to form sodium chloride. Another salt, calcium chloride (which is used to melt ice on roads), has the formula $CaCl_2$; one atom of calcium combines with two atoms of chlorine to form this compound. Carbon tetrachloride, a common compound of chlorine used for dry cleaning, has the formula CCl_4 where C stands for a carbon atom which combines with four chlorine atoms. Another common substance, ammonia, has the formula NH_3; in this case one atom of nitrogen combines with three atoms of hydrogen.

There are especially important examples of combining capacity among the gaseous elements. For example, hydrogen occurs in nature in the form of molecules each of which contains two hydrogen atoms. The molecule of hydrogen consists of two atoms and has the formula H_2. Similarly chlorine has the molecular formula Cl_2. Chemical analysis always gives these results. It would be wrong to try to assign the formula H_3 or H_4 to a molecule of hydrogen, or Cl, Cl_3 or Cl_4 to a molecule of chlorine. These formulas would just not agree with the results of experiments on the composition and properties of hydrogen or chlorine.

The above examples indicate that different elements have different capacities for chemical combination. It was natural for chemists to seek an explanation for these differences. They asked the question: why does a substance have a certain molecular formula and not some other formula? An answer would be possible were we to assume that each species of atom is characterized by some particular combining capacity, or valence. At one time valence was considered as though it might represent the number of hooks possessed by a given atom, and thus the number of links that an atom could form with others of the same or different species. If hydrogen and chlorine atoms each had just one hook (that is, a valence of 1) we would readily understand how it is that molecules like H_2, Cl_2 and HCl are stable, while certain other species like H_3, H_2Cl, HCl_2 and Cl_3 don't exist at all. And if the hydrogen atom is thus assigned a valence of 1, the formula of water (H_2O) requires that the oxygen atom has two hooks or a valence of 2. The formula NH_3 for ammonia leads us to assign a valence of three to nitrogen; the formula CH_4 for methane leads us to assign a valence of 4 to carbon; and so on. Proceeding in this fashion, we can assign a valence number to each of the known elements. Sometimes complications arise as, for example, in the case of sulfur. In H_2S the sulfur atom seems to have a valence of 2, but in such a compound as sulfuric acid (H_2SO_4) sulfur seems to have a valence of 6. In this case and others, then, we may have to assign two (or even more) valence numbers to a single species of atom. At the other extreme of possibilities are those elements, for example, helium, neon and argon, which have not been found as parts of compounds—and to these elements we may appropriately assign a valence of zero.

SG 17.8

The atomic mass and valence are numbers that can be assigned to an element; they are "numerical characterizations" of the atoms of the element. There are other numbers which represent properties of the atoms of the elements, but atomic mass and valence were the two most important to nineteenth-century chemists. These numbers were used in the attempt to find order and regularity among the elements—a problem which will be discussed in the next section.

Q5 At this point we have two numbers which are characteristic of the atoms of an element. What are they?

Q6 Assume the valence of oxygen is 2. In each of the following molecules, give the valence of the atoms other than oxygen: CO, CO_2, NO_3, Na_2O and MnO.

17.4 The search for order and regularity among the elements.

By 1872 sixty-three elements were known; they are listed in Table 17.1 with their atomic masses and chemical symbols. Sixty-three elements are many more than Aristotle's four; and chemists tried to make things simpler by looking for ways of organizing what they had learned about the elements. They tried to find relationships among the elements—a quest somewhat like Kepler's earlier search for rules that would relate the motions of the planets of the solar system.

Summary 17.4
1. Various elements, having similar physical and chemical properties (e.g., valence, density, melting point) can be considered as belonging to a family.
2. The members of a family appear at intervals in a list of elements arranged in order of their atomic mass.

17.4

Relationships did indeed appear: there seemed to be families of elements with similar properties. One such family consists of the so-called alkali metals—lithium, sodium, potassium, rubidium and cesium—listed here in order of increasing atomic mass. We have placed these elements in boxes in Table 17.1. All these metals are similar physically: they are soft and have low melting points. The densities of these metals are very low; in fact, lithium, sodium and potassium are less dense than water. The alkali metals are also similar chemically: they all have valence 1; they all combine with the same elements to form similar compounds. Because they form compounds readily with other elements; they are said to be highly reactive. They do not occur free in nature, but are always found in combination with other elements.

See "Looking for a New Law" in Project Physics Reader 5.

In 1829 the German chemist Johann Wolfgang Döbereiner noticed that elements often formed groups of three members with similar chemical properties. He identified the "triads": chlorine, bromine and iodine; calcium, strontium and barium; sulfur, selenium and tellurium; iron, cobalt and manganese. In each "triad," the atomic mass of the middle member was approximately the arithmetical average of the masses of the other two elements.

Another family of elements, called the halogens, includes, in order of increasing atomic mass, fluorine, chlorine, bromine and iodine. The halogens may be found in Table 17.1 just above the alkali metals, and they have been circled. It turns out that each halogen precedes an alkali metal in the list, although the order of the listing was simply by atomic mass.

Although these four halogen elements exhibit some marked dissimilarities (for example, at 25°C the first two are gases, the third a liquid, the last a volatile solid) they have much in common. They all combine violently with many metals to form white, crystalline salts (halogen means "salt-former") having similar formulas, such as NaF, NaCl, NaBr and NaI, or MgF_2, $MgCl_2$, $MgBr_2$ and MgI_2. From much similar evidence chemists noticed that all four members of the family seem to have the same valence with respect to any other particular element. All four elements form simple compounds with hydrogen (HF, HCl, HBr, HI) which dissolve in water and form acids. All four, under ordinary conditions, exist as diatomic molecules, that is, each molecule contains two atoms.

In 1865 the English chemist J. A. R. Newlands pointed out that the elements could usefully be listed simply in the order of increasing atomic mass. For when this was done, a curious fact became evident: not only were the atomic masses of the elements within any one family regularly spaced, as Döbereiner had suggested, but there was also in the whole list a periodic recurrence of elements with similar properties: "...the eighth element, starting from a given one, is a kind of repetition of the first, like the eighth note in an octave of music." Newlands' proposal was met with skepticism. One chemist even suggested that Newlands might look for similar patterns in an alphabetical list of elements.

The elements which follow the alkali metals in the list also form a family, the one called the alkaline earth family; this family includes beryllium, magnesium, calcium, strontium and barium. Their melting points and densities are higher than those of the alkali metals. The alkaline earths all have a valence of two, and are said to be divalent. They react easily with many elements but not as easily as do the alkali metals.

The existence of these families of elements encouraged chemists to look for a systematic way of arranging the elements so that the members of a family would group together. Many schemes were suggested; the most successful was that of the Russian chemist, D.I. Mendeleev.

Q7 What are three properties of elements which recur systematically with increasing atomic mass?

17.5 Mendeleev's periodic table of the elements.

Mendeleev, examining the properties of the elements, came to the conconclusion that the atomic masses supplied the fundamental "numerical characterization" of the elements. He discovered that if the elements were arranged in a table in the order of their atomic masses—but in a special way—the different families appeared in columns of the table. In his own words:

Although chemically similar elements did occur at periodic intervals, Newlands did not realize that the number of elements in a period changed if one continued far enough. This was recognized by Mendeleev.

> The first attempt which I made in this way was the following: I selected the bodies with the lowest atomic weights and arranged them in the order of the size of their atomic weights. This showed that there existed a period in the properties of the simple bodies, and even in terms of their atomicity the elements followed each other in the order of arithmetic succession of the size of their atoms:
>
> Li=7 Be=9.4 B=11 C=12 N=14 O=16 F=19
> Na=23 Mg=24 Al=27.4 Si=28 P=31 S=32 Cl=35.3
> K=39 Ca=40 . . . . Ti=50 V=51 et cetera

In this table, hydrogen was omitted because of its unique properties. Helium and the other elements of the family of noble gases had not yet been discovered.

Mendeleev set down seven elements, from lithium to fluorine, in the order of increasing atomic masses, and then wrote the next seven, from sodium to chlorine, in the second row. The periodicity of chemical behavior is already evident before we go on to write the third row. In the first vertical column are the first two alkali metals. In the seventh column are the first two halogens. Indeed, within *each* of the columns the elements are chemically similar, having, for example, the same characteristic valence.

Summary 17.5

1. Mendeleev arranged a table of the elements in order of increasing atomic masses. In the table, the known family relationships are clearly evident and the physical and chemical properties show periodic behavior.

2. The tabular arrangement led to the prediction of new elements and the specification of their properties, which were later confirmed by observation.

When Mendeleev added a third row of elements, potassium (K) came below elements Li and Na, which are members of the same family and have the same valence, namely, 1. Next in the row is Ca, divalent like Mg and Be above it. In the next space to the right, the element of next higher atomic mass should appear. Of the elements known at the time, the next heavier was titanium (Ti), and it was placed in this space under Al and B by various workers who had tried to develop such schemes. Mendeleev, however, recognized that Ti has chemical properties similar to those of

Dmitri Ivanovich Mendeleev (men-deh-lay'-ef) (1834-1907) received his first science lessons from a political prisoner who had been banished to Siberia by the Czar. Unable to get into college in Moscow, he was accepted in St. Petersburg, where a friend of his father had some influence. In 1866 he became a professor of chemistry there; in 1869 he published his first table of the sixty-three known elements arranged according to atomic mass. His paper was translated into German at once and was made available to all scientists. Mendeleev came to the United States, where he studied the oil fields of Pennsylvania in order to advise his country on the development of the Caucasian resources.

C and Si and therefore should be put in the fourth vertical column (the pigment, titanium white, TiO_2, has a formula comparable to CO_2 and SiO_2, and all three elements show a valence of 4). Then if the classification is to be complete, there should exist a hitherto unsuspected element with atomic mass between that of Ca (40) and Ti (50) and with a valence of 3. Here was a definite prediction, and Mendeleev found other cases of this sort among the remaining elements.

Table 17.2 is Mendeleev's periodic system or "periodic table" of the elements, proposed in 1872. We note that he distributed the 63 elements then known (with 5 in doubt) in 12 horizontal rows or series, starting with hydrogen at the top left, and ending with uranium at the bottom right. All are written in order of increasing

Table 17.2 Periodic classification of the elements; Mendeleev, 1872.

GROUP→		I	II	III	IV	V	VI	VII	VIII
Higher oxides and hydrides		R_2O —	RO —	R_2O_3 —	RO_2 H_4R	R_2O_5 H_3R	RO_3 H_2R	R_2O_7 HR	RO_4 —
SERIES	1	H(1)							
	2	Li(7)	Be(9.4)	B(11)	C(12)	N(14)	O(16)	F(19)	
	3	Na(23)	Mg(24)	Al(27.3)	Si(28)	P(31)	S(32)	Cl(35.5)	
	4	K(39)	Ca(40)	—(44)	Ti(48)	V(51)	Cr(52)	Mn(55)	Fe(56), Co(59), Ni(59), Cu(63)
	5	[Cu(63)]	Zn(65)	—(68)	—(72)	As(75)	Se(78)	Br(80)	
	6	Rb(85)	Sr(87)	?Yt(88)	Zr(90)	Nb(94)	Mo(96)	—(100)	Ru(104), Rh(104), Pd(106), Ag(108)
	7	[Ag(108)]	Cd(112)	In(113)	Sn(118)	Sb(122)	Te(125)	I(127)	
	8	Cs(133)	Ba(137)	?Di(138)	?Ce(140)	—	—	—	
	9	—	—	—	—	—	—	—	
	10	—	—	?Er(178)	?La(180)	Ta(182)	W(184)	—	Os(195), Ir(197), Pt(198), Au(199)
	11	[Au(199)]	Hg(200)	Tl(204)	Pb(207)	Bi(208)	—	—	
	12	—	—	—	Th(231)	—	U(240)		

atomic mass (Mendeleev's values given in parentheses), but are so placed that elements with similar chemical properties are in the same vertical column or group. Thus in Group VII are all the halogens; in Group VIII, only metals that can easily be drawn to form wires; in Groups I and II, metals of low densities and melting points; and in Group I, the family of alkali metals.

Anyone particularly interested in Mendeleev can be directed to his original paper (in English) in the World of Mathematics, II, pp. 913-918.

Table 17.2 shows many gaps. But, as Mendeleev realized, it revealed an important generalization:

> For a true comprehension of the matter it is very important to see that all aspects of the distribution of the elements according to the order of their atomic weights express essentially one and the same fundamental dependence—periodic properties.

By this is meant that in addition to the gradual change in physical and chemical properties within each vertical group, there is also a periodic change of properties in the horizontal sequence, beginning with hydrogen and ending with uranium.

This periodic law is the heart of the matter. We can best illustrate it as Lothar Meyer did, by drawing a curve showing the values of some physical quantity as a function of atomic mass. Figure 17.1 is a plot of the atomic volumes of the elements. This atomic volume is defined as the atomic mass of the substance divided by its density in the liquid or solid state. Each circled point on this graph represents an element; a few of the points have been labeled with the identifying chemical symbols. Viewed as a whole, the graph demonstrates a striking periodicity: as the mass increases the atomic volume first drops, then

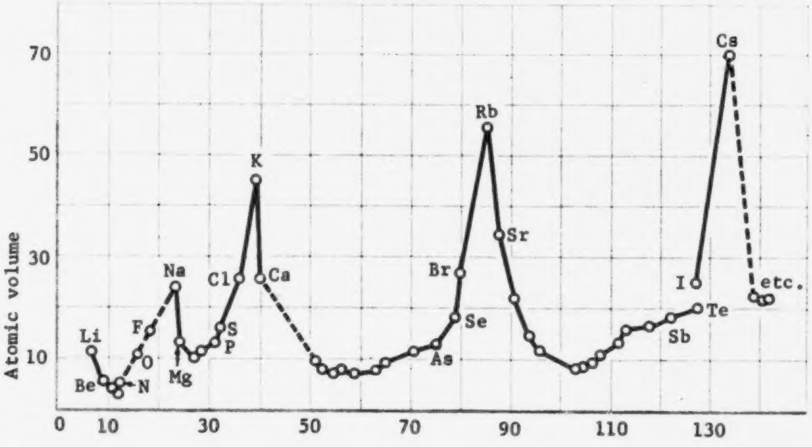


Fig. 17.1 The atomic volumes of elements graphed against their atomic masses

17.5

increases to a sharp maximum, drops off again and increases to another sharp maximum, and so on. And at the successive peaks we find Li, Na, K, Rb, Cs, the members of the family of alkali metals. On the left-hand side of each peak, there is one of the halogens.

In 1864, the German chemist Lothar Meyer wrote a chemistry textbook. In this book, he considered how the properties of the chemical elements might depend on their atomic masses. He later found that if he plotted the atomic volume against the atomic mass, the line drawn through the plotted points rose and fell in two short periods, then in two long periods. This was exactly what Mendeleev had discovered in connection with valence. Mendeleev published his result in 1869; Meyer published his in 1870. Meyer, as he himself later admitted, lacked the courage to predict the discovery of unknown elements. Nevertheless, Meyer should be given part of the credit for the idea of the periodic table.

Mendeleev's periodic table of the elements not only provided a remarkable correlation of the elements and their properties, it also enabled him to predict that certain unknown elements must exist and what many of their properties should be. To estimate physical properties of a missing element, Mendeleev averaged the properties of its nearest neighbors in the table: those to right and left, above and below. A striking example of Mendeleev's success in using the table in this way is his set of predictions concerning the gap in Series 5 Group IV. This was a gap in Group IV, which contained elements with properties resembling those of carbon and silicon. Mendeleev assigned the name "eka-silicon" (Es) to the unknown element. His predictions of the properties of this element are listed in the left-hand column that follows. In 1887, this element was isolated and identified (it is now called "germanium"); its properties are listed in the right-hand column.

"The following are the properties which this element should have on the basis of the known properties of silicon, tin, zinc, and arsenic.	The predictions in the left column were made by Mendeleev in 1871. In 1887 an element (germanium) was discovered which was found to have the following properties:
Its atomic [mass] is nearly 72,	Its atomic mass is 72.5.
it forms a higher oxide EsO_2,...	It forms an oxide GeO_2, and
Es gives volatile organo-metallic compounds; for instance...$Es(C_2H_5)_4$,	forms an organo-metallic compound $Ge(C_2H_5)_4$
which boil at about 160°, etc.;	which boils at 160° C, and
also a volatile and liquid chloride, $EsCl_4$, boiling at about 90°	forms a liquid chloride $GeCl_4$ which boils at 83° C
and of specific gravity about 1.9....	and has a specific gravity of 1.9.
the specific gravity of Es will be about 5.5, and EsO_2 will have a specific gravity of about 4.7, etc...."	The specific gravity of germanium is 5.5 and the specific gravity of GeO_2 is 4.7.

Mendeleev's predictions are remarkably close to the properties actually found.

The daring of Mendeleev is shown in his willingness to venture detailed numerical predictions; the sweep and power of his system is shown above in the remarkable accuracy of those predictions. In similar fashion, Mendeleev described the properties to be expected for the then unknown elements in Group III, Period 4 and in Group III, Period 5, elements now called gallium and scandium, and again his predictions turned out to be remarkably accurate.

Even though not every aspect of Mendeleev's work yielded such successes, these were indeed impressive results. Successful numerical predictions like these are among the most desired results in physical science.

Q8 What was the basic ordering principle in Mendeleev's table?

Q9 What reasons led him to violate that principle?

Q10 How did he justify leaving gaps in the table?

17.6 The modern periodic table. The periodic table has had an important place in chemistry and physics for nearly one hundred years. It presented a serious challenge to any theory of the atom proposed after about 1880: the challenge of providing an explanation for the order among the elements expressed by the table. A successful model of the atom must provide a physical explanation for the of the elements. In Chapter 19 we shall see how one model of the atom—the Bohr model—met this challenge.

Since 1872 many changes have had to be made in the periodic table, but they have been changes in detail rather than in general ideas. None of these changes has affected the basic feature of periodicity among the properties of the elements. A modern form of the table is shown in Table 17.3.

Summary 17.6

1. The modern periodic table accommodates forty additional elements including a new family, the noble gases, and two new series, the lanthanides and the actinides. With these additions the periodicity among properties remains intact.

2. A basic change from the Mendeleev arrangement is the ordering by atomic number rather than atomic mass.

Table 17.3 A modern form of the periodic table of the elements. The number above the symbol is the chemist's atomic weight, the number below the symbol is the atomic number.

Group→ Period↓	I	II											III	IV	V	VI	VII	0
1	1.0080 H 1																	4.0026 He 2
2	6.939 Li 3	9.012 Be 4											10.811 B 5	12.011 C 6	14.007 N 7	15.999 O 8	18.998 F 9	20.183 Ne 10
3	22.990 Na 11	24.31 Mg 12											26.98 Al 13	28.09 Si 14	30.97 P 15	32.06 S 16	35.45 Cl 17	39.95 Ar 18
4	39.10 K 19	40.08 Ca 20	44.96 Sc 21	47.90 Ti 22	50.94 V 23	52.00 Cr 24	54.94 Mn 25	55.85 Fe 26	58.93 Co 27	58.71 Ni 28	63.54 Cu 29	65.37 Zn 30	69.72 Ga 31	72.59 Ge 32	74.92 As 33	78.96 Se 34	79.91 Br 35	83.80 Kr 36
5	85.47 Rb 37	87.62 Sr 38	88.91 Y 39	91.22 Zr 40	92.91 Nb 41	95.94 Mo 42	(99) Tc 43	101.07 Ru 44	102.91 Rh 45	106.4 Pd 46	107.87 Ag 47	112.40 Cd 48	114.82 In 49	118.69 Sn 50	121.75 Sb 51	127.60 Te 52	126.9 I 53	131.30 Xe 54
6	132.91 Cs 55	137.34 Ba 56	* 57–71	178.49 Hf 72	180.95 Ta 73	183.85 W 74	186.2 Re 75	190.2 Os 76	192.2 Ir 77	195.09 Pt 78	196.97 Au 79	200.59 Hg 80	204.37 Tl 81	207.19 Pb 82	208.98 Bi 83	210 Po 84	(210) At 85	222 Rn 86
7	(223) Fr 87	226.05 Ra 88	† 89															

*Rare-earth metals	138.91 La 57	140.12 Ce 58	140.91 Pr 59	144.27 Nd 60	(147) Pm 61	150.35 Sm 62	151.96 Eu 63	157.25 Gd 64	158.92 Tb 65	162.50 Dy 66	164.93 Ho 67	167.26 Er 68	168.93 Tm 69	173.04 Yb 70	174.97 Lu 71
† Actinide metals	227 Ac 89	232.04 Th 90	231 Pa 91	238.03 U 92	(237) Np 93	(242) Pu 94	(243) Am 95	(245) Cm 96	(249) Bk 97	(249) Cf 98	(253) E 99	(255) Fm 100	(256) Mv 101	(253) No 102	(257) Lw 103

17.6

One difference between the modern and older tables is that new elements have been added. Forty new elements have been identified since 1872, so that the table now contains 103 elements. Some of these new elements are especially interesting, and we shall need to know something about them.

Although Mendeleev's table had eight columns, the column labelled VIII did not contain a family of elements. It contained the "transition" elements which are now in the long series (periods) labelled 4,5 and 6 in Table 17.3. The group labelled "0" in Table 17.3 does consist of a family of elements, the noble gases, which do have similar properties.

Comparison of the modern form of the table with Mendeleev's table shows that the modern table contains eight groups, or families, instead of seven. The additional group is labeled "zero." In 1894, the British scientists Lord Rayleigh and William Ramsay discovered that about 1 percent of our atmosphere consists of a gas that had previously escaped detection. It was given the name argon (symbol Ar). Argon does not seem to enter into chemical combination with any other elements, and is not similar to any of the groups of elements in Mendeleev's original table. Other elements similar to argon were also discovered: helium (He), neon (Ne), krypton (Kr), xenon (Xe), and radon (Rn). These elements are considered to form a new group or family of elements, called the "noble gases." (In chemistry, elements such as gold and silver that react only rarely with other elements were called "noble" and all the members of the new family are gases at room temperature.) Each noble gas (with the exception of argon) has an atomic mass slightly smaller than that of a Group I element. The molecules of the noble gases contain only one atom, and until only a few years ago no compound of any noble gas was known. The group number zero was thought to correspond to the chemical inertness, or zero valence of the members of the group. In 1963, some compounds of xenon and krypton were produced, so that these elements are not really inert. These compounds are not found in nature, and are unstable when they are made in the laboratory. The noble gases are certainly less liable to react chemically than any other elements and their position in the table does correspond to their "reluctance" to react.

Helium was first detected in the spectrum of the sun in 1868 (Chapter 19). Its name comes from helios, the Greek word for the sun. It was not discovered on earth until 1895, when Ramsay found it in a uranium-containing mineral (Chapter 21). Almost all the helium in the world comes from natural gas wells in Texas, Kansas and Oklahoma. Helium is lighter than air, and is widely used in balloons and blimps instead of highly flammable hydrogen.

T35: Periodic table

In addition to the noble gases, two other sets of elements had to be included in the table. After the fifty-seventh place, room had to be made for a whole set of 14 chemically almost indistinguishable elements, known as the rare earths or lanthanide series. Most of these elements were unknown in Mendeleev's time. Similarly, a set of 14 very similar elements, forming what is called the actinide series, belongs immediately after actinium at the eighty-ninth place. These elements are shown in two rows below the main table. No more additions are now expected within the table. There are no known gaps, and we shall see in Chapters 19 and 20 that according to the best theory of the atom now available, no new gaps should appear.

Besides the addition of new elements to the periodic table, there have also been some changes of a more general type. As we have seen, Mendeleev arranged the elements in order of increasing atomic mass. In the late nineteenth century, however, this basic scheme was found to break down in several places. For example, the chemical properties of argon (Ar) and potassium (K) demand that they should be placed in the eighteenth and nineteenth positions, whereas on the basis of their atomic masses alone (39.948 for argon, 39.102 for potassium), their positions should be reversed. Other reversals of this kind have been found necessary, for example, for the fifty-second element, tellurium (at. mass = 127.60) and the fifty-third, iodine (at. mass = 126.90). The consecutive integers that indicate the number for the best position for the element, according to its chemical properties, are called the atomic numbers; the atomic number is usually denoted by the symbol Z; thus for hydrogen, Z = 1; for uranium, Z = 92. The atomic numbers of all the elements are given in Table 17.3. In Chapter 19 we shall see that the atomic number has a fundamental physical meaning related to atomic structure.

The need for reversals in the periodic table of the elements would have been a real catastrophe• to Mendeleev. He confidently expected, for example, that the atomic mass of tellurium (modern value = 127.60, fifty-second place), when more accurately determined, would turn out to be lower than that of iodine (modern value = 126.90, fifty-third place) and, in fact, in 1872 (see Table 17.2) he had convinced himself that the correct atomic mass of tellurium was 125! Mendeleev overestimated the necessity of the periodic law in every detail, particularly as it had not yet received a physical explanation. Although the reversals in the sequence of elements have proved to be real (e.g., tellurium, in fifty-second place, does have a higher atomic mass than iodine, in fifty-third place in the periodic table), their existence did not invalidate the scheme. Satisfactory explanations for these reversals have been found in modern atomic physics.

• It would have been very disturbing, but perhaps catastrophe is too strong a word; few theories fit all the data perfectly.

SG 17.9
SG 17.10
SG 17.11

Q11 What is the "atomic number" of an element?

Summary 17.7
The invention of the electric battery made possible the study of the response of chemical compounds to electric current. Compounds that had previously defied decomposition could easily be decomposed. Such 'electrolysis' led to the discovery of several new elements.

17.7 Electricity and matter: qualitative studies. While chemists were applying Dalton's atomic theory, another development was taking place which opened an important path to our understanding of the atom. Sir Humphry Davy and Michael Faraday made discoveries which showed that electricity and matter are intimately related. Their work marked the beginning of electrochemistry. Their discoveries had to do with the breaking down, or decomposition, of chemical compounds by electric currents. This process is called electrolysis.

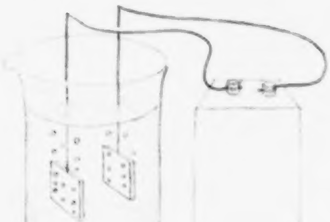
The study of electrolysis was made possible by the invention of the electric cell by the Italian scientist Alessandro Volta, in 1800. Volta's cell consisted of a pair of zinc and copper discs, separated from each other by a sheet of paper moistened with a weak salt solution. As a result of chemical changes occurring in the cell, an electric potential difference is established across the cell. A battery usually consists of several similar cells connected together.

A battery has two terminals, one positively charged and the other negatively charged. When the terminals are connected to each other, outside the battery, by means of certain materials, there is an electric current in the battery and the materials. We say that we have a circuit. The connecting materials in which the current exists are called conductors of electricity. Thus, the battery can produce and maintain an electric current. It is not the only device that can do so, but it was the first source of steady currents.

Not all substances are electrical conductors. Among solids, the metals are the best conductors. Some liquids conduct electricity. Pure distilled water is a poor conductor. But when certain substances such as acids or salt are dissolved in water, the resulting solutions are good electrical conductors. Gases are not conductors under normal conditions, but can be made electrically conducting in the presence of strong electric fields, or by other methods. The conduction of electricity in gases, vital to the story of the atom, will be discussed in Chapter 18.

Within a few weeks after Volta's announcement of his discovery it was found that water could be decomposed into oxygen and hydrogen by the use of electric currents. Figure 17.2 is a diagram of an electrolysis apparatus. The two terminals of the battery are connected, by conducting wires, to two thin sheets of platinum. When

Fig. 17.2



E41*: Electrolysis effect

D53: Electrolysis of water

these platinum sheets are immersed in ordinary water, bubbles of oxygen appear at one sheet and bubbles of hydrogen at the other. Adding a small amount of certain acids speeds up the reaction without changing the products. Hydrogen and oxygen gases are formed in the proportion of 7.94 grams of oxygen to 1 gram of hydrogen, which is exactly the proportion in which these elements combine to form water. Water had previously been impossible to decompose, and had been regarded—from ancient times until after 1750—as an element. Thus the ease with which water was separated into its elements by electrolysis dramatized the chemical use of electricity, and stimulated many other investigations of electrolysis.

Among these investigations, some of the most successful were those of the young English chemist Humphry Davy. Perhaps the most striking of Davy's successes were those he achieved when, in 1807, he studied the effect of the current from a large electric battery on soda and potash. Soda and potash were materials of commercial importance (for example, in the manufacture of glass, soap and gunpowder) and had been completely resistant to every earlier attempt to decompose them. Soda and potash were thus regarded as true chemical elements—up to the time of Davy's work. When electrodes connected to a powerful battery were touched to a solid lump of soda, or to a lump of potash, part of the solid was heated to its melting point. At one electrode gaseous oxygen was released violently; at the other electrode small globules of molten metal appeared which burned brightly and almost explosively in air. When the electrolysis was done in the absence of air, the metallic material could be obtained. Sodium and potassium were discovered in this way. The metallic element sodium was obtained from soda (in which it is combined as sodium hydroxide) and the metallic element potassium obtained from potash (in which it is combined as potassium hydroxide). In the immediately succeeding years electrolytic trials made on several hitherto undecomposed "earths" yielded the first samples ever obtained of such metallic elements as magnesium, strontium and barium: there were also many other demonstrations of the striking changes produced by the chemical activity of electricity.

Q12 Why was the first electrolysis of water such a surprising phenomenon?

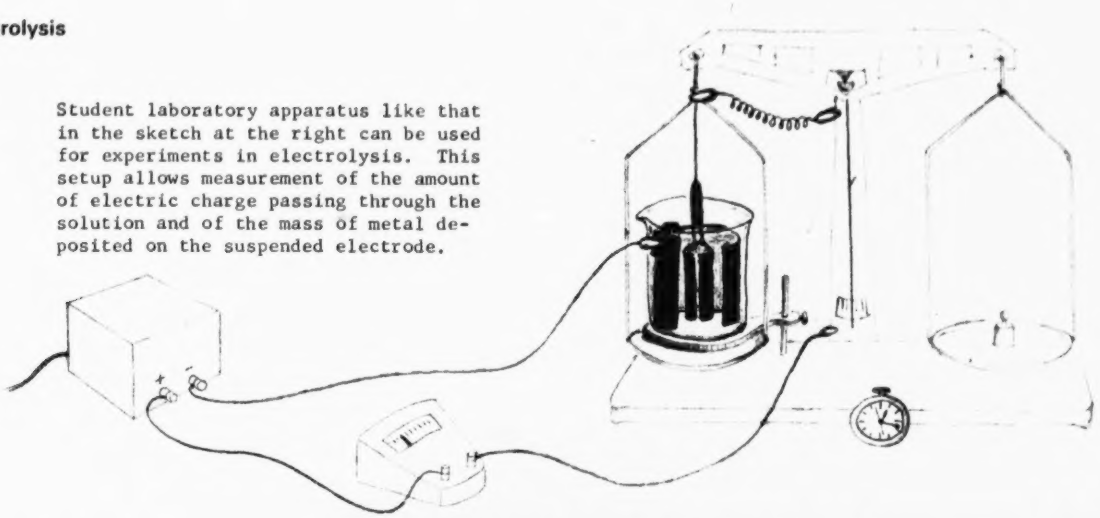
Q13 Some equally striking results of electrolysis followed. What were they?

F37: Counting electrical charges in motion

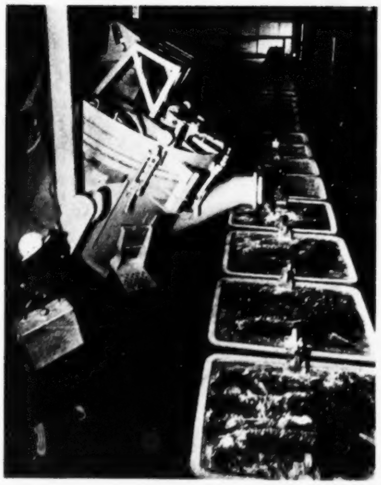
This can be explained by assuming that some of the water molecules come apart, leaving the hydrogen atoms with a + charge and the oxygen atoms with a - charge; the hydrogen atoms would be attracted to the - plate and the oxygen to the + plate. Faraday called the charged atoms ions (after the Greek word for "wanderers"). Solutions of such charged particles are said to be ionized.

Humphry Davy (1778-1829) was the son of a farmer. In his youth he worked as an assistant to a physician but was discharged because of his liking for explosive chemical experiments. He became a chemist, discovered nitrous oxide (laughing gas), later used as an anaesthetic, and developed a safety lamp for miners. His work in electrochemistry and his discovery of several elements made him world-famous; he was knighted in 1812. In 1813 Sir Humphry Davy hired a young man, Michael Faraday, as his assistant and took him along on an extensive trip through France and Italy. It became evident to Davy that young Faraday was a man of scientific genius. Davy is said to have been envious, at first, of Faraday's great gifts. He later said that he believed his greatest discovery was Faraday.

Electrolysis



Student laboratory apparatus like that in the sketch at the right can be used for experiments in electrolysis. This setup allows measurement of the amount of electric charge passing through the solution and of the mass of metal deposited on the suspended electrode.



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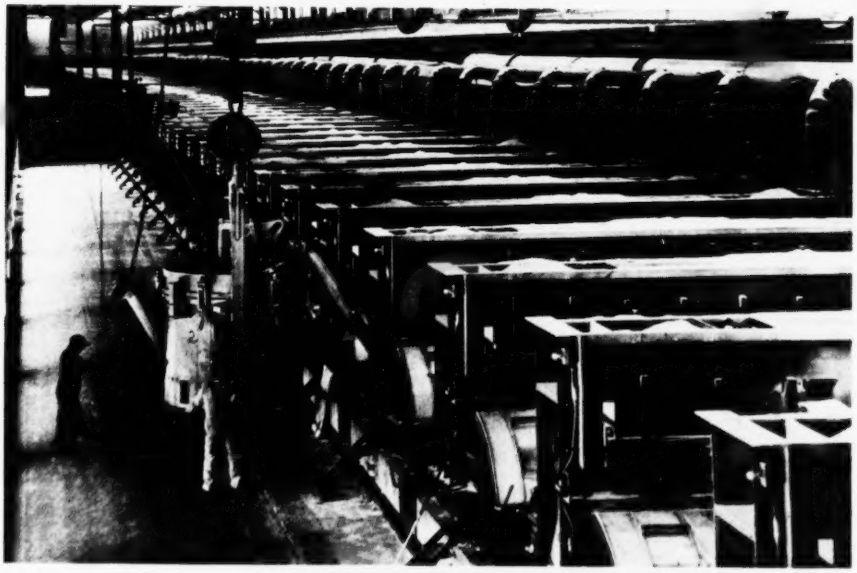
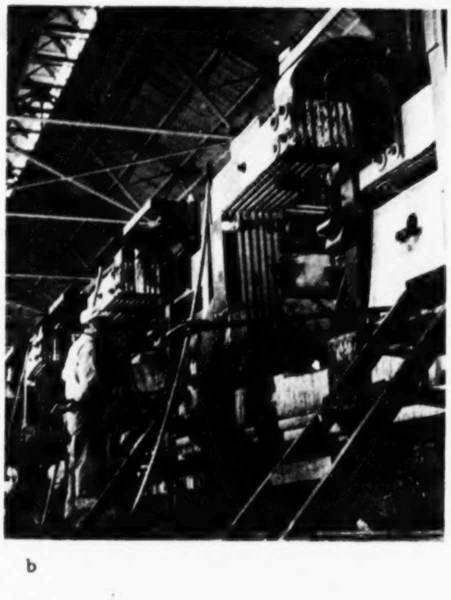


a

The separation of elements by electrolysis is important in industry, particularly in the production of aluminum. These photographs show the vast scale of a plant where aluminum is separated out of aluminum ore in electrolytic tanks.

- a) A row of tanks where aluminum is separated out of aluminum ore.
- b) A closer view of the front of some tanks, showing the thick copper straps that carry the current.
- c) A huge vat of molten aluminum that has been siphoned out of the tanks is poured into molds.



b

17.8 Electricity and matter: quantitative studies. Davy's work on electrolysis was mainly qualitative. But quantitative questions were also asked. How much chemical change can be produced by a given amount of electricity? If solutions of different chemical compounds are electrolyzed with a given amount of current, how do the amounts of chemical change produced compare? Will doubling the amount of electricity double the chemical effects?

Summary 17.8
Electrochemical studies showed that the chemical change produced by electrolysis is related to the total charge passing through the cell, to the atomic mass of the elements involved, and to their valences. Thus electric charge has a precise quantitative relationship to atomic properties.

Answers to these questions were supplied by Michael Faraday, who discovered two fundamental laws of electrolysis. He studied the electrolysis of a solution of copper sulfate, a blue salt, in water. He made an electrolytic cell by immersing two bars of copper in the solution and attaching them to the terminals of a battery. The electric current that flowed through the resulting circuit caused copper from the solution to be deposited on the cathode and oxygen to be liberated at the anode. Faraday determined the amount of copper deposited by weighing the anode before the electrolysis started and again after a known amount of current had passed through the solution. He measured the current with an ammeter. Faraday found that the mass of copper deposited depends on two things: on the magnitude (say, in amperes) of the current (I), and on the length of time (t) that the current was maintained. In fact, the mass of copper deposited was directly proportional to both the current and the time. When either was doubled, the mass of copper deposited was doubled. When both were doubled, four times as much copper was deposited. Similar results were found in experiments on the electrolysis of many different substances.

Faraday's results may be described by stating that the amount of chemical change produced in electrolysis is proportional to the product: $I \cdot t$. Now, the current (in amperes) is the quantity of charge (in coulombs) which moves through the electrolytic cell per unit time (in seconds). The product $I \cdot t$ therefore gives the total charge that has moved through the cell during the given experiment. We then have Faraday's first law of electrolysis:

> The mass of electrolytically liberated chemicals is proportional to the amount of charge which has passed through the electrolytic cell.

Next Faraday measured the amounts of different elements liberated from chemical compounds by given amounts of electric charge, that is, by different values of the product I·t. He found that the amount of an element produced by a given amount of electricity depends on the atomic mass of the element and on the valence of the element. His second law of electrolysis states:

> If A is the atomic mass of an element, and if v is its valence, a certain amount of electric charge, 96,540 coulombs, produces A/v grams of the element.

This amount of electric charge, 96,540 coulombs, is called one faraday.

Table 17.4 gives examples of Faraday's second law of electrolysis.

Table 17.4. Masses of elements produced from compounds by 96,540 coulombs of electric charge.

Element	Atomic Mass	Valence	Mass of Element Produced (grams)
Hydrogen	1.008	1	1.008
Chlorine	35.45	1	35.45
Oxygen	16.00	2	8.00
Copper	63.54	2	31.77
Zinc	65.37	2	32.69
Aluminum	26.98	3	8.99

SG 17.12
SG 17.13
SG 17.14
SG 17.15

The mass of the element produced is seen to be equal to the atomic mass divided by the valence. This quantity, A/v, is a measure of the amount of one element that combines with another element. For example, the ratio of the amounts of oxygen and hydrogen liberated by 96,540 coulombs of electric charge is $\frac{8.00}{1.008} = 7.94$. But this is just the ratio of the mass of oxygen to the mass of hydrogen in water.

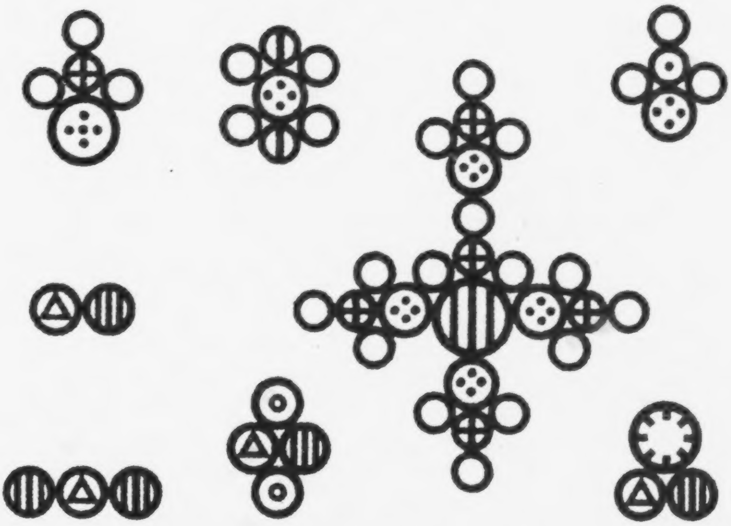
Faraday's second law of electrolysis has an important implication. A given amount of electric charge is somehow closely connected with the atomic mass and valence of an element. The mass and valence are characteristic of the atoms of the element. Perhaps, then, a certain amount of electricity is somehow connected with an atom of the element. The implication is that electricity may also be atomic in nature. This possibility was considered by Faraday, who wrote:

> ...if we adopt the atomic theory or phraseology, then the atoms of bodies which are equivalents to each other in their ordinary chemical action have equal quantities of electricity naturally associated with them. But I must confess that I am jealous of the term atom; for though it is very easy to talk of atoms, it is very difficult to form a clear idea of their nature, especially when compound bodies are under consideration.

In Chapter 18 you will read about the details of the research that established the atomic nature of electricity. This research was of great and fundamental importance, and helped make possible the exploration of the structure of the atom.

SG 17.16
SG 17.17
SG 17.18

Q14 The amount of an element deposited in electrolysis depends on three factors. What are they?



Dalton's visualization of the composition of various compounds

80.3% zinc. 19.7% oxygen

17.1 The chemical compound zinc oxide (molecular formula ZnO) contains equal numbers of atoms of zinc and oxygen. Using values of atomic masses from the modern version of the periodic table, find the percentage by mass of zinc in zinc oxide. What is the percentage of oxygen in zinc oxide?

17.2 The chemical compound zinc chloride (molecular formula $ZnCl_2$) contains two atoms of chlorine for each atom of zinc. Using values of atomic masses from the modern version of the periodic table, find the percentage by mass of zinc in zinc chloride. 47.9%

17.3 From the decomposition of a 5.00-gram sample of ammonia gas into its component elements, nitrogen and hydrogen, 4.11 grams of nitrogen were obtained. The molecular formula of ammonia is NH_3. Find the mass of a nitrogen atom relative to that of a hydrogen atom. Compare your result with the one you would get by using the values of the atomic masses in the modern version of the periodic table. If your result is different from the latter result, how do you account for the difference? N/H = 13.9/1

17.4 From the information in Problem 17.3, calculate how much nitrogen and hydrogen are needed to make 1.2 kg of ammonia. 986 g of N and 214 g of H

17.5 If the molecular formula of ammonia were NH_2, and you used the result of the experiment of Problem 17.3, what value would you get for the ratio of the mass of a nitrogen atom relative to that of a hydrogen atom? N/H = 9.23/1

17.6 A sample of nitric oxide gas, weighing 1.00 g, after separation into its components, is found to have contained 0.47 g of nitrogen. Taking the atomic mass of oxygen to be 16.00, find the corresponding numbers that express the atomic mass of nitrogen relative to oxygen on the respective assumptions that the molecular formula of nitric oxide is (a) NO; (b) NO_2; (c) N_2O. (a) 14.1 (b) 28.2 (c) 7.0

17.7 Early data yielded 8/9.2 for the mass ratio of nitrogen and oxygen atoms, and 1/7 for the mass ratio of hydrogen and oxygen atoms. Show that these results lead to a value of 6 for the relative atomic mass of nitrogen, provided that the value 1 is assigned to hydrogen. discussion

17.8 Given the molecular formulae HCl, NaCl, $CaCl_2$, $AlCl_3$, $SnCl_4$, PCl_5, find possible valence numbers of sodium, calcium, aluminum, tin and phosphorus. Na = 1, Ca = 2, Al = 3, Sn = 4, P = 5

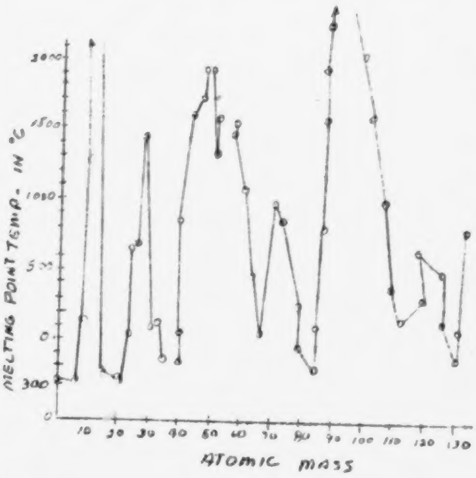
17.9 a) Examine the modern periodic table of elements and cite all reversals of order of increasing atomic mass. 7 reversals

b) Restate the periodic law in your own words, not forgetting about these reversals. discussion

17.10 In recent editions of the Handbook of Chemistry and Physics there are printed in or below one of the periodic tables the valence numbers of the elements. Neglect the negative valence numbers and plot (to element 65) a graph of maximum valences observed vs. atomic mass. What periodicity is found? Is there any physical or chemical significance to this periodicity? Does there have to be any? See next page

17.11 Look up the data in the Handbook of Chemistry and Physics, then plot some other physical characteristic against the atomic masses of the elements from hydrogen to barium in the periodic table. Comment on the periodicity (melting point, boiling point, etc.).

17.11



17.12 According to Table 17.4, when 96,500 coulombs of charge pass through a water solution, 1.008 g of hydrogen and how much of oxygen will be released? How much hydrogen and how much oxygen will be produced when a current of 3 amperes is passed through water for 60 minutes (3600 seconds)?

0.113 g hydrogen
0.895 g oxygen

17.13 If a current of 0.5 amperes is passed through molten zinc chloride in an electrolytic apparatus, what mass of zinc will be deposited in

a) 5 minutes (300 seconds); 0.05 g
b) 30 minutes; 0.30 g
c) 120 minutes? 1.20 g

17.14 a) For 20 minutes (1200 seconds) a current of 2.0 amperes is passed through molten zinc chloride in an electrolytic apparatus. What mass of chlorine will be released at the anode? 0.88 g Cl

b) If the current had been passed through molten zinc iodide rather than molten zinc chloride what mass of iodine would have been released at the anode? 3.4 g

c) Would the quantity of zinc deposited in part (b) have been different from what it was in part (a)? Why? No; Faraday's first law

17.15 How is Faraday's speculation about an "atom of electricity" related to atomicity in the chemical elements? discussion

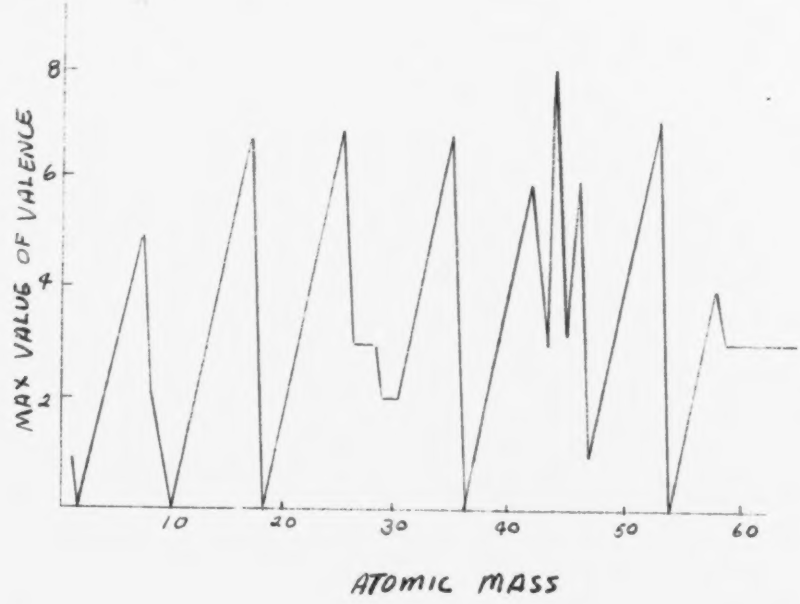
17.16 The idea of chemical elements composed of identical atoms makes it easier to correlate the phenomena discussed in this chapter. Could the phenomena be explained without using the idea of atoms? Are chemical phenomena, which usually involve a fairly large quantity of material (in terms of the number of "atoms" involved), sufficient evidence for belief in the atomic character of materials? discussion

17.17 A sociologist recently wrote a book about the place of man in modern society called Multivalent Man. In what sense might he have used the term "multivalent?" discussion

17.18 Compare the atomic theory of the Greeks, as described in the prologue to this chapter, with the atomic theories described in Unit 3. (You will probably need to consult reference books for more details of the theory. The best reference is probably Lucretius, On the Nature of Things.)

discussion

#17.10



Chapter 18 Electrons and Quanta

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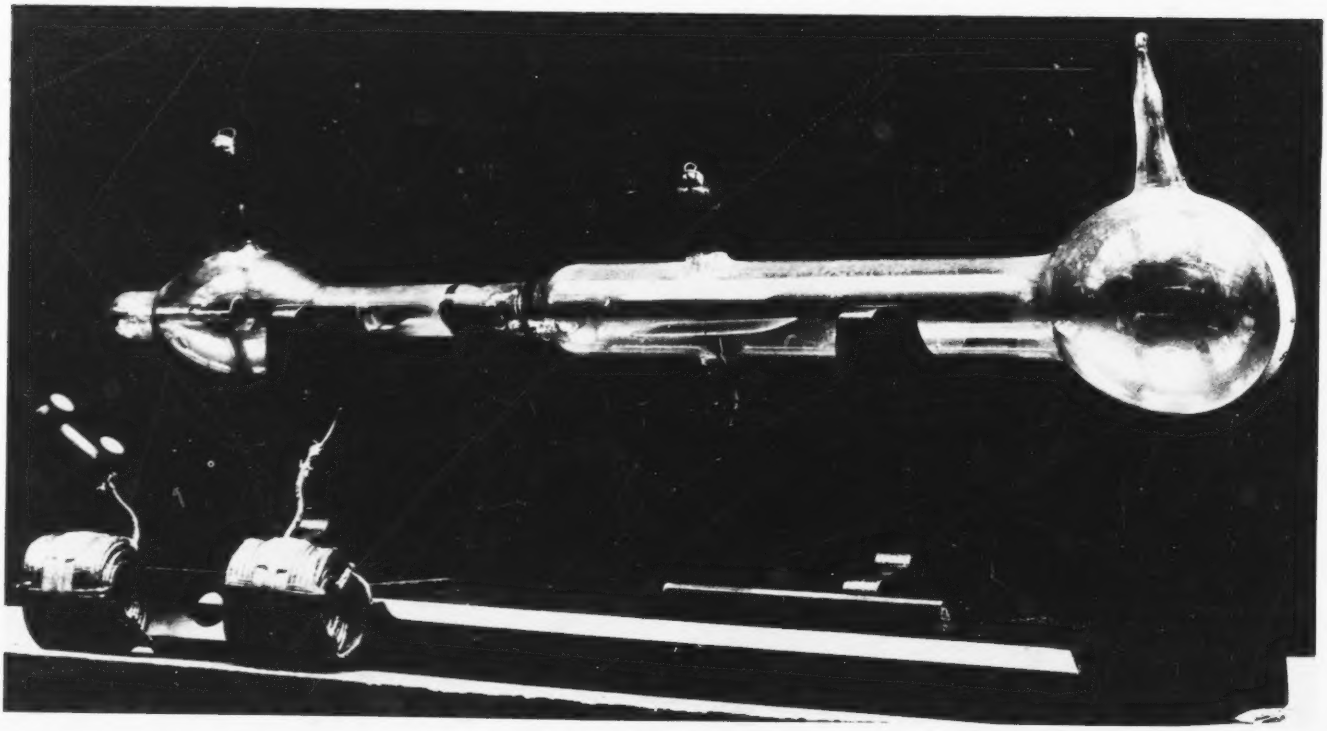
The tube used by J.J. Thomson to determine the charge to mass ratio of electrons.



"The electron has conquered physics and many worship the new idol rather blindly"
Poincaré, 1907.

18.1 The problem of atomic structure: pieces of atoms.

The development of chemistry in the nineteenth century raised the general question: are atoms really indivisible, or do they consist of still smaller particles? We can see the way in which this question arose by thinking a little more about the periodic table. Mendeleev had arranged the elements in the order of increasing atomic mass. But the atomic masses of the elements cannot explain the periodic features of Mendeleev's table. Why, for example, do the third, eleventh, nineteenth, thirty-seventh, fifty-fifth and eighty-seventh elements, with quite different atomic masses, have similar chemical properties? Why are these properties somewhat different from those of the fourth twelfth, twentieth, thirty-eighth, fifty-sixth and eighty-eighth elements in the list, but greatly different from the properties of the second, tenth, eighteenth, thirty-sixth, fifty-fourth and eighty-sixth elements? The differences in atomic mass were not enough to account, by themselves, for the differences in the properties of the elements. Other reasons had to be sought.

These elements burn when exposed to air; they decompose water, often explosively.

These elements combine slowly with air and water.

These elements rarely combine with anything.

The periodicity of the properties of the elements led to speculation about the possibility that atoms might have structure, that they might be made up of smaller pieces. The gradual changes of properties from group to group might suggest that some unit of atomic structure is added, in successive elements, until a certain portion of the structure is completed. The completed condition might occur in a noble gas. In atoms of the next heavier element, a new portion of the structure may be started, and so on. The methods and techniques of classical chemistry could not supply experimental evidence for such structure. In the nineteenth century, however, discoveries and new techniques in physics opened the way to the proof that atoms do, indeed, consist of smaller pieces. Evidence piled up which showed that the atoms of different elements differ in the number and arrangement of these pieces, or building blocks.

Summary 18.1

1. The periodic features of the elements could not be explained on the basis of atomic masses.

2. The periodicity of the properties of the elements suggests that atoms have a structure which builds up in some regular way.

3. The studies of newly discovered phenomena were combined to affirm that atoms have an inner structure.

In this chapter, we shall discuss the discovery of one kind of piece which atoms contain: the electron. We shall see how experiments with light and electrons led to a revolutionary idea—that light energy is transmitted in discrete amounts. In Chapter 19, we shall describe the discovery of another part of the atom, the nucleus. Then we shall show how Niels Bohr combined these pieces to create a workable model of the atom. The story starts with the discovery of cathode rays.

E 39: The charge-to-mass ratio for an electron

D54: Charge-to-mass ratio for an electron

18.2 Cathode rays. In 1855 a German physicist, Heinrich Geissler, invented a vacuum pump which could remove enough gas from a strong glass tube to reduce the pressure to 0.01 percent of normal air pressure. His friend, Julius Plücker, connected one of Geissler's evacuated tubes to a battery. He was surprised to find that at the very low gas pressure that could be obtained with Geissler's pump, electricity flowed through the tube. Plücker used apparatus similar to that shown in Fig. 18.1. He sealed a wire into each end of a strong glass tube. Inside the tube, each wire ended in a metal plate, called an electrode. Outside the tube, each wire ran to a source of high voltage. (The negative plate is called the cathode, and the positive plate is called the anode.) A meter indicated the current in the tube.

Geissler (1814-1879) made the first major improvement in vacuum pumps after Guericke invented the air pump two centuries earlier.

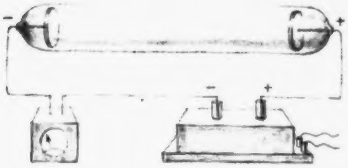


Fig. 18.1 Cathode ray apparatus.

Plücker and his student, Johann Hittorf, noticed that when an electric current passes through a tube at low gas pressure, the tube itself glows with a pale green color. Plücker described these effects in a paper published in 1858. He wrote:

> ...a pale green light...appeared to form a thin coating immediately upon the surface of the glass bulb.... the idea forcibly presented itself that it was a fluorescence in the glass itself. Nevertheless the light in question is in the inside of the tube; but it is situated so closely to its sides as to follow exactly [the shape of the tubes], and thus to give the impression of belonging to the glass itself.

Substances which glow when exposed to light are called fluorescent. Fluorescent lights are essentially Geissler tubes with an inner coating of fluorescent powder.

Several other scientists observed these effects, but two decades passed before anyone undertook a thorough study of the glowing tubes. By 1875, Sir William Crookes had designed new tubes for studying the glow produced when an electric current passes through an evacuated tube. When he used a bent tube, as in Fig. 18.2, the most intense green glow appeared on the part of the tube which was directly opposite the cathode. This suggested that the green glow was produced by something which comes out of the cathode and travels straight down the tube until it hits the glass. Another physicist, Eugen Goldstein, who was studying the effects of passing an electric current through a gas at low pressure, named whatever was coming from the cathode, cathode rays.

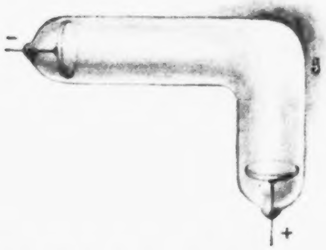


Fig. 18.2 Bent Geissler tube. The most intense green glow appeared at g.

To study the nature of the rays, Crookes did some ingenious experiments. He reasoned that if the cathode rays could be stopped before they reached the end of the tube, the intense green glow should disappear. He therefore introduced a barrier in the form of a Maltese cross, as in Fig. 18.3. Instead of the intense green glow, a shadow of the cross appeared at the end of the tube. The cathode seemed to act like a candle which produces light; the cross acted like a barrier blocking the light. Because the shadow, cross and cathode were lined up, Crookes concluded that the cathode rays, like light rays, travel in straight lines. Next, Crookes moved a magnet near the tube, and the shadow moved. Thus he found that magnetic fields deflected the paths of cathode rays. In the course of many experiments, Crookes found the following properties of cathode rays:

a) cathodes of many different materials produce rays with the same properties;
b) in the absence of a magnetic field, the rays travel in straight lines perpendicular to the surface that emits them;
c) a magnetic field deflects the path of the cathode rays;
d) the rays can produce chemical reactions similar to the reactions produced by light: for example, certain silver salts change color when hit by the rays.

Crookes suspected, but did not succeed in showing that

e) charged objects deflect the path of cathode rays.

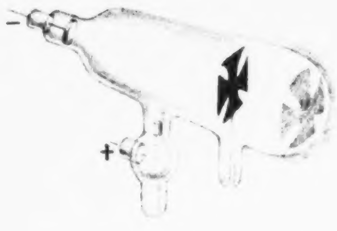


Fig. 18.3 Crookes' tube.

J. J. Thomson observed this in 1897.

Physicists were fascinated by the cathode rays and worked hard to understand their nature. Some thought that the rays must be a form of light, because they have so many of the properties of light: they travel in straight lines, produce chemical changes and fluorescent glows just as light does. According to Maxwell's theory of electricity and magnetism (Unit 4) light consists of electromagnetic waves. So the cathode rays might be electromagnetic waves of frequency higher or lower than that of visible light.

Magnetic fields, however, do not bend light; they do bend the path of cathode rays. In Unit 4, we found that magnets exert forces on currents, that is, on moving electric charges. Since a magnet deflects cathode rays in the same way that it deflects negative charges, some physicists believed that cathode rays consisted of negatively charged particles.

The controversy over the wave or particle nature of cathode rays continued for 25 years. Finally, in 1897, J. J. Thomson made a series of experiments which convinced physicists that the cathode rays are negatively charged particles.

It was known that the paths of charged particles are affected by both magnetic and electric fields. By assuming

Summary 18.2

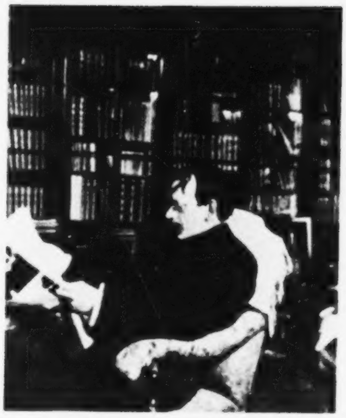
1. Gases at low pressure conduct electricity. In a gas discharge tube, "rays" appear to come from the negative terminal (cathode).

2. Thomson established that cathode rays are a stream of negatively-charged particles, by measuring their charge-to-mass ratio q/m

3. The surprisingly large value of q/m led Thomson to conclude that the particles have a very small mass compared to atoms.

4. Because identical particles were obtained from many different gases, Thomson suggested that they might be fundamental building blocks of atoms.

that the cathode rays were negatively charged particles, Thomson could predict what should happen to the cathode rays when they passed through such fields. For example, the deflection of the path of the cathode rays by a magnetic field could be just balanced by an electric field with the right direction and magnitude. The predictions were verified and Thomson could conclude that the cathode rays did indeed act like charged particles. He was then able to calculate, from the experimental data, the ratio of the charge of a particle to its mass. This ratio is denoted by q/m where q is the charge and m is the mass of the particle. For those who are interested in the details of Thomson's experiment and calculations, they are given on page 41.



Sir Joseph John Thomson (1856-1940), one of the greatest British physicists, attended Owens College in Manchester, England (home of John Dalton) and then Cambridge University. Throughout his career, he was interested in atomic structure. We shall read about his work often during the rest of the course. He worked on the conduction of electricity through gases, on the relation between electricity and matter and on atomic models. His greatest single contribution was the discovery of the electron. He was the head of the famous Cavendish Laboratory at Cambridge University, where one of his students was Ernest Rutherford—about whom you will hear a great deal later in this unit and in Unit 6.

Thomson found that the rays from cathodes made of different materials all had the same value of q/m, namely 1.76×10^{11} coulombs per kilogram. This value was about 1800 times larger than the values of q/m for hydrogen ions measured in electrolysis experiments, 9.6×10^{7} coulombs per kilogram. Thomson concluded from these results that either the charge of the cathode ray particles was much larger than that of the hydrogen ion, or the mass of the cathode ray particles was much smaller than the mass of the hydrogen ion.

Thomson's negatively charged particles were later called electrons. Thomson also made measurements of the charge on the negatively charged particles with methods other than those involving deflection by electric and magnetic fields. Although these experiments were inaccurate, they were good enough to indicate that the charge of a cathode ray particle was not much different from that of the hydrogen ion in electrolysis. Thomson was therefore able to conclude that the cathode ray particles have much smaller mass than hydrogen ions.

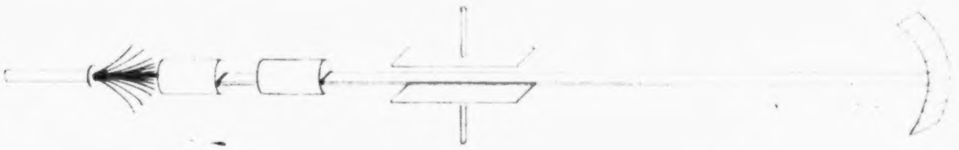
SG 18.1

The cathode ray particles, or electrons, were found to have two important properties: (1) they were emitted by a wide variety of cathode materials, and (2) they were much smaller in mass than the hydrogen atom, which has the smallest known mass. Thomson therefore concluded that the cathode ray particles form a part of all kinds of matter. He suggested that the atom is not the ultimate limit to the subdivision of matter, and that the electron is one of the bricks of which atoms are built up, perhaps even the fundamental building block of atoms.

The name "electron", for the fundamental unit of electricity was introduced in 1874 by the Irish physicist G. Johnstone Stoney (1826-1911).

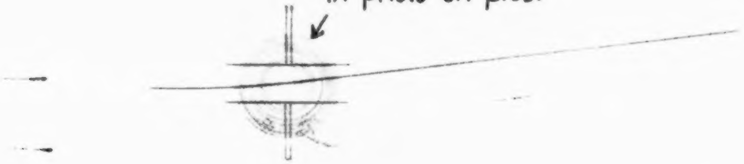
Thomson's q/m Experiment

J. J. Thomson measured the ratio of charge to mass for cathode-ray particles by means of the evacuated tube in the photograph on page 36. A high voltage between two electrodes in one end of the tube produced cathode rays. The rays that passed through both slotted cylinders formed a nearly parallel beam. The beam produced a spot of light on a fluorescent coating inside the large end of the tube.



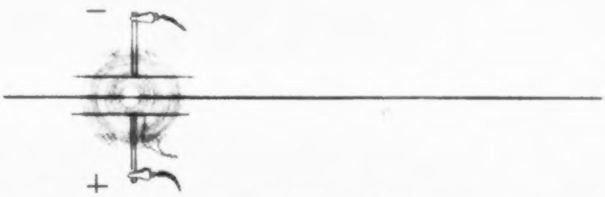
The beam could be deflected by an electric field produced between two plates in the mid-section of the tube.

The beam could also be deflected by a magnetic field produced between a pair of wire coils placed around the mid-section of the tube.



relate this to coil in photo on p.36.

When only the magnetic field B was turned on, the particles in the beam of charge q and speed v would experience a force Bqv; because the force is always perpendicular to the velocity, the beam would be deflected into a nearly circular arc of radius R in the nearly uniform field. If the particles in the beam have mass m, they must be experiencing a centripetal force mv^2/R. Since the centripetal force is the magnetic force, $Bqv = mv^2R$. Rearranging terms: $q/m = v/BR$.
B can be calculated from the geometry of the coils and the current in them. R can be found geometrically from the displacement of the beam spot on the end of the tube. To determine v, Thomson applied the electric field and the magnetic field at the same time. By arranging the directions and strengths of the fields appropriately, the electric field can be made to exert a downward force Eq on the beam particles exactly equal to the upward force Bqv due to the magnetic field.



If the magnitudes of the electric and magnetic forces are equal, then $Eq = Bqv$. Solving for v: $v = E/B$. E can be calculated from the separation of the two plates and the voltage between them, so the speed of the particles can be determined. So all the terms on the right of the equation for q/m are known and q/m can be found.

18.2

The year 1897 may be considered as marking the end of the age-old concept of the atom as indivisible.

In the article in which he announced his discovery, Thomson speculated on the ways in which the particles could be arranged in atoms of different elements in order to account for the periodicity of the chemical properties of the elements. Although, as we shall see, he did not say the last word about this problem, he did say the first word about it.

Q1 What was the most convincing evidence that cathode rays were not electromagnetic radiation?

Q2 Why was q/m for electrons 1800 times larger than q/m for hydrogen ions?

Q3 What were two main reasons that Thomson believed electrons to be "building blocks" from which all atoms are made?

Summary 18.3
1. By measuring the electric force on tiny charged oil droplets, Millikan showed that there is a smallest charge of which all others are multiples.

2. By combining this value for q with Thomson's value of q/m, the mass m of the unit of charge, the "electron", can be found.

18.3 The measurement of the charge of the electron: Millikan's experiment. After the ratio of charge to the mass (q/m) of the electron had been determined, physicists tried to measure the value of the charge q separately. If the charge could be determined, the mass of the electron could be found from the known value of q/m. In the years between 1909 and 1916 an American physicist, Robert A. Millikan, succeeded in measuring the charge of the electron. This quantity is one of the fundamental constants of physics because of its importance in atomic and nuclear physics as well as in electricity and electromagnetism.

Millikan's "oil-drop experiment" is described on page 43. He found that the electric charge that an oil drop picks up is always a simple multiple of a certain minimum value. For example, the charge may have the value 4.8×10^{-19} coulombs, or 1.6×10^{-19} coulombs, or 6.4×10^{-19} coulombs, or 1.6×10^{-18} coulombs. But it never has a charge of, say, 2.4×10^{-19} coulombs, and it never has a value smaller than 1.6×10^{-19} coulombs. In other words, electric charges always come in multiples of 1.6×10^{-19} coulombs. Millikan took this minimum charge to be the charge of a single electron.

From now on we denote the magnitude of the charge of the electron by q_e:

$q_e = 1.6 \times 10^{-19}$ coul.

SG 18.2

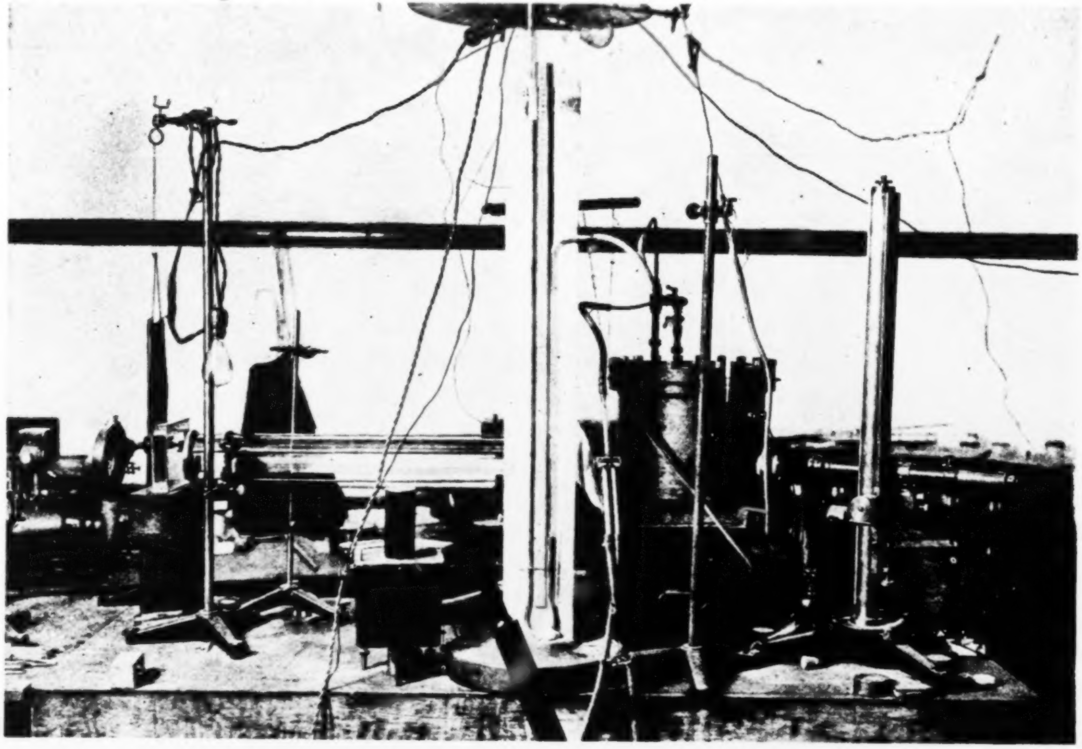
E40*: The measurement of elementary charge.

Charges of atomic and molecular ions are measured in units of the electron charge q_e. For example, when a chemist refers to a "doubly charged oxygen ion," he means that the charge of the ion is $2q_e = 3.2 \times 10^{-19}$ coulombs.

Note that Millikan's experiments did not prove that no smaller charges than q_e can exist. All we can say is that no experiment has yet proved the existence of smaller charges. Since Millikan's work, physicists have been convinced that electric charges always come in multiples of q_e.

In 1897, the British physicist Townsend obtained a value of the ionic charge of approximately 1×10^{-19} coulomb using Stokes' relations and an early Wilson cloud chamber.

In 1903, H.A. Wilson, making measurements on a cloud of water droplets between charged horizontal plates, obtained the average charge of a droplet to vary from 0.7×10^{-19} to 1.4×10^{-19} coulomb.



Millikan's Oil-Drop Experiment

In principle Millikan's experiment is simple; it is sketched in Fig. 18.5. When oil is sprayed into a chamber, the minute droplets formed are found to be electrically charged. The charge on a droplet can be measured by means of an electric field in the chamber. Consider a small oil drop of mass m carrying an electric charge q. It is situated between two horizontal plates separated by a distance d and at an electrical potential difference V. There will be a uniform electric field $\vec{E}$ between the plates, of strength V/d. This field can be adjusted so that the electrical force $q\vec{E}$ exerted upward on the drop's charge will balance the force $m\vec{g}$ exerted downward by gravity. Equating the magnitudes of these forces gives:

$$F_{el} = F_{grav},$$

$$qE = mg,$$

$$\text{or} \qquad q = mg/E.$$

The mass of the drop can, in principle, be determined from its radius and the density of the oil from which it was made. Millikan had to measure these quantities by an indirect method, but it is now possible to do the experiment with small manufactured polystyrene spheres whose mass is accurately known, so that some of the complications of the original experiment can be avoided. Millikan's own set-up is seen in the photograph above. A student version of Millikan's apparatus is shown in the photograph at the right.

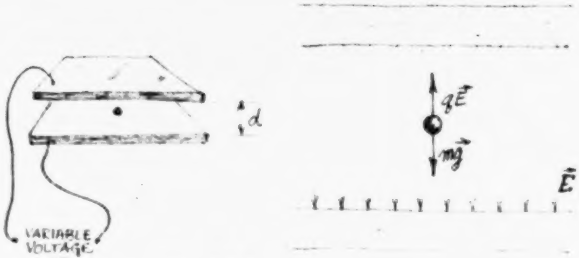
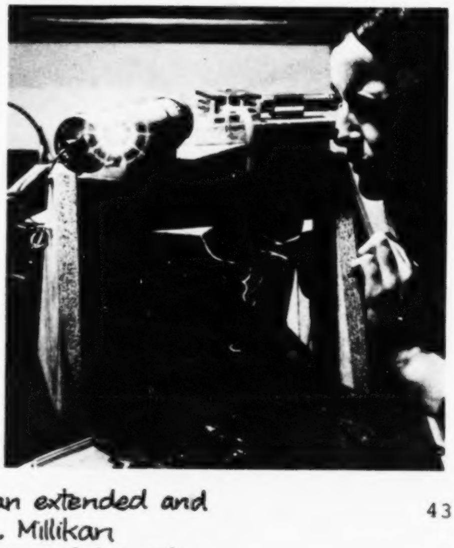


Fig. 18.5



Wilson's problem was twofold: (1) making measurements on an extended and amorphous cloud, and (2) the evaporation of the droplets. Millikan removed the first by observing a single droplet and the second by using oil instead of water.

Having measured q_e (Sec. 18.3), q_e/m (Sec. 18.2) and the F (Sec. 17.8), Avogadro's number (N_o) as well as the masses of the electron (m_e) and of the hydrogen atom (m_H) may all be obtained.

18.3

In 1964, an American physicist, Murray Gell-Mann, suggested that particles with charge equal to 1/3 or 2/3 of q_e might exist. He named these particles "quarks"—the word comes from James Joyce's novel Finnegan's Wake. Quarks are now being looked for in cosmic-ray and bubble-chamber experiments.

In everyday life, the electric charges one meets are so large that one can think of a current as being continuous—just as one usually thinks of the flow of water in a river as continuous rather than as a flow of individual molecules. A current of one ampere, for example, is equivalent to the flow of 6.25×10^{18} electrons per second. The "static" electric charge one accumulates by shuffling over a rug on a dry day consists of approximately 10^{12} electron charges.

Best values at present are:
$m_e = .91085 \times 10^{-30}$ kg
$m_H = 1.6724 \times 10^{-27}$ kg
$N_o = 6.0247 \times 10^{23}$

Since the work of Millikan, a wide variety of other experiments involving many different fields within physics have all pointed to the same basic unit of charge as being fundamental in the structure and behavior of atoms. For example, it has been shown directly that cathode ray particles carry this basic unit of charge—that they are, in other words, electrons.

By combining Millikan's value for the electron charge q_e with Thomson's value for the ratio of charge to mass (q_e/m), we can calculate the mass of a single electron (see margin). The result is that the mass of the electron is about 10^{-30} kilograms. The charge-to-mass ratio of a hydrogen ion is 1836 times smaller than the charge-to-mass ratio of an electron. It is reasonable to consider that an electron and a hydrogen ion have equal and opposite electric charge, since they form a neutral hydrogen atom when they combine. We may therefore conclude that the mass of the hydrogen ion is 1836 times as great as the mass of the electron.

Thomson found that

$$\frac{q_e}{m} = 1.76 \times 10^{11} \text{ coul/kg.}$$

According to Millikan's experiment the magnitude of q_e is

1.6×10^{-19} coul.

Therefore, the mass of an electron is:

$$m = \frac{1.6 \times 10^{-19} \text{ coul}}{1.76 \times 10^{11} \text{ coul/kg}}$$

$$= 0.91 \times 10^{-30} \text{ kg.}$$

(Mass of a hydrogen ion is 1.66×10^{-27} kg. This is approximately the value of one "atomic mass unit.")

Q4 Oil drops pick up different amounts of electric charge. How did Millikan know that the lowest charge he found was actually just one electron charge?

Summary 18.4

1. The experimental results associated with the photoelectric effect could not be understood in terms of the classical electromagnetic theory of light.

2. The experimental results causing the chief difficulty were

a) the existence of a threshold frequency,

b) the short time required for electron ejection, and

c) the relation between electron energy and light frequency.

18.4 The photoelectric effect. The photoelectric effect was discovered in 1887 by the German physicist Heinrich Hertz. Hertz was testing Maxwell's theory of electromagnetic waves (Unit 4). He noticed that a metallic surface can emit electric charges when light of very short wave length falls on it. Because light and electricity are both involved, the name photoelectric effect was given to this new phenomenon. When the electricity produced was passed through electric and magnetic fields, its direction was changed in the same ways as the path of cathode rays. It was therefore deduced that the electricity consists of negatively charged particles. In 1898, J. J. Thomson measured the value of the ratio q/m for these particles with the same method that he used for the cathode ray particles. He got the same value for the particles ejected in the photoelectric effect as he did for the cathode ray particles. By means of these experiments (and others) the photoelectric particles were shown to have the same properties as electrons. They are often referred to as *photoelectrons* to indicate their source. Later work showed that all substances, solids, liquids and gases, undergo the photoelectric effect under appropriate conditions. It is, however, convenient to study the effect with metallic surfaces.

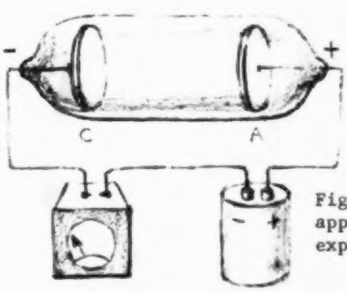
The "electric eye" used, for example, for opening a door automatically, is based on the photoelectric effect. When a solid object interrupts a beam of light shining from one side of the door to the other, an electric current is changed; this change switches on a motor that operates the door. The photoelectric effect is also used in projectors for sound motion pictures.

The photoelectric effect has been studied in great detail and has had an important place in the development of atomic physics. The effect could not be explained in terms of the classical physics we have studied so far. New ideas had to be introduced to account for the experimental results. In particular, a revolutionary concept had to be introduced—that of *quanta*—and a new branch of physics—*quantum theory*—had to be developed, at least in part because of the photoelectric effect. Modern atomic theory is actually the quantum theory of matter and radiation. The study of the photoelectric effect is, therefore, an important step on the way to the understanding of the atom.

In 1888, the German physicist Hallwachs was the first to observe that a negatively charged zinc plate lost its charge when it was illuminated with ultra-violet light.

Two types of measurements can be made which yield useful information about the photoelectric effect: (1) measurements of the photoelectric current (the number of electrons emitted per unit time); (2) measurements of the kinetic energies with which the electrons are emitted.

The electron current can be studied with an apparatus like that sketched in Fig. 18.6. Two metal plates, C and A, are sealed inside a well-evacuated quartz tube. (Quartz is transparent to ultraviolet light as well as visible light.) The two plates are connected to a source of potential difference. When light strikes plate C, electrons are emitted. If the potential of plate A is positive relative to plate C, the emitted electrons will accelerate to plate A. (Some electrons will reach plate A even if it isn't positive relative to C.) The resulting current is indicated by the meter.



E42*: Photoelectric effect

D55: Photoelectric effect

Fig. 18.6a Schematic diagram of apparatus for photoelectric experiments.

The term "electric eye" was coined at the 1933 Chicago World's Fair.

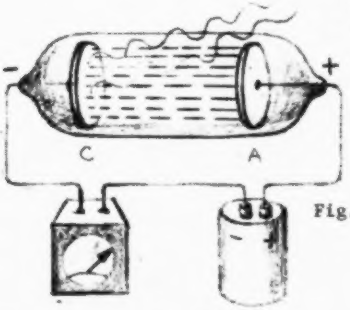


Fig. 18.6b

T36: Photoelectric mechanism

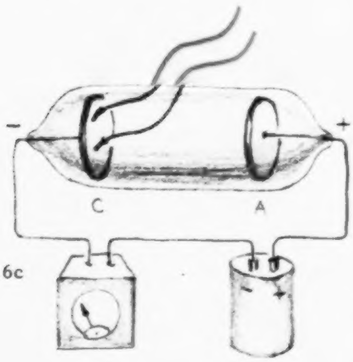


Fig. 18.6c

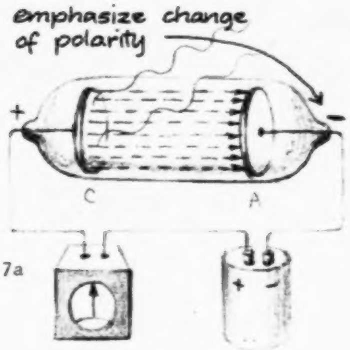

emphasize change of polarity

Fig. 18.7a

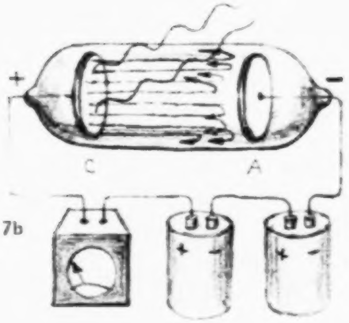


Fig. 18.7b

Any metal used as the plate C shows a photoelectric effect, but only if the light has a frequency greater than a certain value. This value of the frequency is called the threshold frequency. Different metals have different threshold frequencies. If the incident light has a frequency lower than the threshold frequency, no electrons are emitted no matter how great the intensity of the light is or how long the light is left on.

The kinetic energies of the electrons can be measured in a slightly modified version of the apparatus of Fig. 18.6. The battery is reversed so that the plate A repels the electrons. The voltage can be varied from zero to a value just large enough to keep any electrons from reaching the plate A. A sketch of the modified apparatus is shown in Fig. 18.7.

SG 18.2

When the voltage across the plates is zero, the meter indicates a current, showing that the electrons emerge from the metallic surface with kinetic energy. As the voltage is increased the electron current decreases until a certain voltage is reached at which the current becomes zero. This voltage, which is called the stopping voltage, is a measure of the maximum kinetic energy of the photoelectrons. If the stopping voltage is denoted by V_{stop}, the maximum kinetic energy is given by the relation:

$$KE_{max} = \frac{1}{2}mv^2_{max} = V_{stop}\,q_e.$$

The results may be stated more precisely. (We shall number the important experimental results to make it more convenient to discuss their theoretical interpretation later.)

(1) A substance shows a photoelectric effect only if the incident radiation has a frequency above a certain value called the threshold frequency.

(2) If light of a given frequency can liberate electrons from a surface, the current is proportional to the intensity of the light.

(3) If light of a given frequency can liberate electrons, the emission of the electrons is immediate. The time interval between the incidence of the light on the metallic surface and the appearance of electrons is not more than 3×10^{-9} sec. This is true even for the lowest light intensities used.

See TG p.46, for derivation of time delay required by wave model.

(4) The maximum kinetic energy of the photoelectrons increases linearly with the frequency of the light which causes their emission, and is independent of the intensity of the incident light. The way in which the maximum kinetic energy of the electrons varies with the frequency of the light is shown in Fig. 18.8. The symbols $(f_0)_1$, $(f_0)_2$ and $(f_0)_3$ stand for the different threshold frequencies of three different substances. For each substance, the experiments fall on a straight line. All the lines have the same slope.

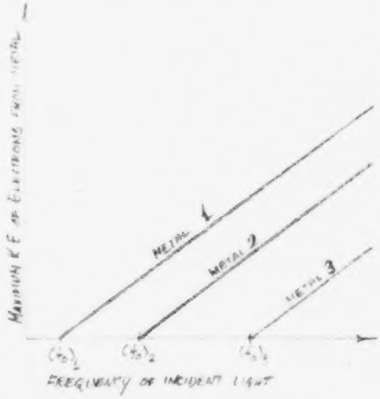


Fig. 18.8 Photoelectric effect: maximum kinetic energy of the electrons as a function of the frequency of the incident light; different metals yield lines that are parallel, but have different threshold frequencies.

What is most surprising about the results is that photoelectrons are emitted at light frequencies barely above the threshold frequency, no matter how low the intensity of the light. Yet, at light frequencies just a bit below the threshold frequency, no electrons are emitted no matter how high the intensity of the light.

With the alkali metals (Li, Na, K, etc.), the photoelectric effect can be produced with visible light.

The experimental results could not be explained on the basis of the classical electromagnetic theory of light. There was no way in which a very low-intensity train of light waves spread out over a large number of atoms could, in a very short time interval, concentrate enough energy on one electron to knock the electron out of the metal. In some experiments. the light intensity was so low that, according to the classical theory, it should take several hundred seconds for an electron to accumulate enough energy from the light to be emitted. But experimental result (3) shows that electrons are emitted about a billionth of a second after the light strikes the surface.

Furthermore, the classical wave theory was unable to account for the existence of a threshold frequency. There seemed to be no reason why a sufficiently intense beam of low-frequency radiation would not be able to produce photoelectricity, if low-intensity radiation of higher frequency could produce it. Finally, the classical theory was unable to account for the fact that the maximum kinetic energy of the photoelectrons increases linearly with the frequency of the light but is independent of the intensity. Thus, the photoelectric effect posed a challenge which the classical wave theory of light could not meet.

Q5 Light falling on a certain metal surface causes electrons to be emitted. What happens as the intensity of the light is decreased?

Q6 What happens as the frequency of the light is decreased?

Summary 18.5
1. To explain the photoelectric effect, Einstein extended Planck's idea of the quantum of thermal radiation to light.

See the articles "Einstein" and "Einstein and some Civilized Discontents" in Project Physics Reader 5.

2. Einstein proposed that light consists of wave pockets or quanta, whose energy is proportional to the frequency of the wave.

3. Einstein's theory was verified by Millikan's experimental results.

$h = 6.6 \times 10^{-34}$ joule-sec

4. It appeared that two models of light, the wave model and the quantum model, were required to account for all the phenomena of light.

In 1902, the German physicist Lenard made the astonishing discovery that the maximum speed of the photoelectrons was independent of the intensity of incident light.

18.5 Einstein's theory of the photoelectric effect: quanta. The explanation of the photoelectric effect was the major work cited in the award to Albert Einstein of the Nobel Prize in physics for the year 1921. Einstein's theory, proposed in 1905, played a major role in the development of atomic physics. The theory was based on a daring proposal. Not only were many of the experimental details unknown in 1905, but the key point of Einstein's explanation was contrary to the classical ideas of the time.

Einstein assumed that the energy of light is not distributed evenly over the whole expanding wave front (as is assumed in the classical theory) but rather is concentrated into discrete small regions. Further, the amount of energy in each of these regions is not just any amount, but is a definite amount of energy which is proportional to the frequency f of the wave. The proportionality factor is a constant, denoted by h and called Planck's constant, for reasons which will be discussed later. Thus, on this model, the light energy comes in pieces, each of amount hf. The amount of radiant energy in each piece is called a quantum of energy; it represents the smallest quantity of energy of light of that frequency. The quantum of light energy was later called a photon.

There is no explanation clearer or more direct than Einstein's. We quote from his first paper (1905) on this subject, changing only the notation used there to make it coincide with usual current practice (including our own notation):

> ...According to the idea that the incident light consists of quanta with energy hf, the ejection of cathode rays by light can be understood in the following way. Energy quanta penetrate the surface layer of the

T37: Photoelectric equation

F39: Photoelectric effect

body, and their energy is converted, at least in part, into kinetic energy of electrons. The simplest picture is that a light quantum gives up all its energy to a single electron; we shall assume that this happens. The possibility is not to be excluded, however, that electrons receive their energy only in part from the light quantum. An electron provided with kinetic energy inside the body may have lost part of its kinetic energy by the time it reaches the surface. In addition it is to be assumed that each electron, in leaving the body, has to do an amount of work W (which is characteristic of the body). The electrons ejected directly from the surface and at right angles to it will have the greatest velocities perpendicular to the surface. The kinetic energy of such an electron is

SG 18.4

$$KE_{max} = hf - W$$

If the body is charged to a positive potential V_{stop} just large enough to keep the body from losing electric charge, we must have

$$KE_{max} = hf - W = V_{stop}\, q_e$$

where q_e is the magnitude of the electronic charge. ...

If the derived formula is correct, V_{stop}, when plotted as a function of the frequency of the incident light, should yield a straight line whose slope should be independent of the nature of the substance illuminated.

SG 18.5
SG 18.6
SG 18.7

We can now compare Einstein's photoelectric equation with the experimental results to test whether or not the theory accounts for the results. According to the equation, the kinetic energy is greater than zero only when the frequency f is high enough so that hf is greater than W. Hence, the equation says that an electron can be emitted only when the frequency of the incident light is greater than a certain value.

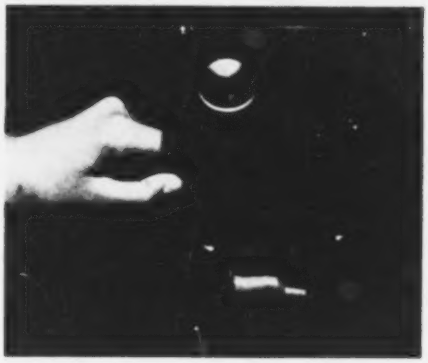
How Einstein's theory explains the photoelectric effect:

(1) no photoelectric emission below threshold frequency. Reason: low-frequency photons don't have enough energy.

(2) current $\propto$ light intensity. Reason: one photon ejects one electron.

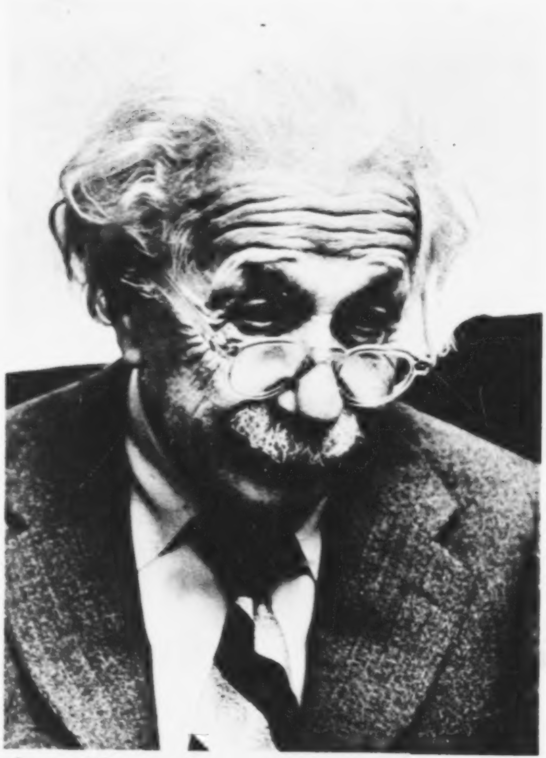
Next, according to Einstein's photon model, it is an individual photon that ejects an electron. The intensity of the light is proportional to the number of the photons, and the number of electrons ejected is proportional to the number of photons. Hence the number of electrons ejected is proportional to the intensity of the incident light.

SG 18.8
SG 18.9



Student apparatus for photoelectric experiments often includes a vacuum phototube like the one at the right (actual size). The collecting wire is at the center of a cylindrical photosensitive surface. The frequency of the light entering the tube is controlled by placing colored filters between the tube and a light source.

Because of Einstein's Jewish ancestry, Marie Curie had to write a letter on his behalf to obtain a professorship in Germany, even after his special theory of relativity was published.



Albert Einstein (1879-1955) was born in the city of Ulm, in Germany. He received his early education in Germany and Switzerland. Like Newton he showed no particular intellectual promise as a youngster. After graduation from the Polytechnic School, Einstein (in 1901) went to work in the Swiss Patent Office in Berne. This job gave Einstein a salary to live on and an opportunity to use his spare time in thinking about physics. In 1905 he published three papers of epoch-making importance. One dealt with quantum theory and included his theory of the photoelectric effect. Another treated the problem of molecular motions and sizes, and worked out a mathematical analysis of the phenomenon of "Brownian motion." Einstein's analysis and experimental work by Jean Perrin, a French physicist, made a strong argument for the molecular motions assumed in the kinetic theory. Einstein's third 1905 paper discusses the theory of special relativity which revolutionized modern thought about the nature of space, time and physical theory.

Later in life Einstein wrote, "If my relativity theories prove right, the Germans will say I am German, the French that I am Swiss, ... the British that I am European .... If they prove wrong, the Germans will say I am a Swiss Jew, the French that I am a German, and the British won't say anything."

In 1915, Einstein published his paper on the theory of general relativity, in which he provided a new theory of gravitation which included Newton's theory as a special case.

When Hitler and the Nazis came to power in Germany, in 1933, Einstein came to the United States and became a member of the Institute for Advanced Studies at Princeton. He spent the rest of his working life seeking a unified theory which would include gravitation and electromagnetics. At the beginning of World War II, Einstein wrote a letter to President Franklin D. Roosevelt warning of the war potential of an "atomic bomb," on which the Germans had begun to work. After World War II, Einstein worked for a world agreement to end the threat of atomic warfare.

According to Einstein's model the light energy is concentrated in the quanta (or photons). Hence, no time is needed for collecting light energy; the quanta transfer their energy immediately to the photoelectrons, which appear after the very short time required for them to escape from the surface.

(3) immediate emission.
Reason: a single photon provides the energy concentrated in one place.

Finally, according to the photoelectric equation, the greater the frequency of the incident light, the greater is the maximum kinetic energy of the ejected electrons. According to the photon model, the photon energy is directly proportional to the light frequency. The minimum energy needed to eject an electron is that required to supply the energy of escape from the metal surface—which explains why light of frequency less than f_o cannot eject any electrons. The difference in the energy of the absorbed photon and the energy lost by the electron in passing through the surface is the kinetic energy of the escaping electron.

(4) KE_{max} increases linearly with frequency above f_o.
Reason: the work needed to remove the electron is $W = hf_o$; any energy left over from the original photon is now available for kinetic energy of the electron.

Thus, Einstein's photoelectric equation agrees qualitatively with the experimental results. There remained two quantitative tests: (1) does the maximum energy vary linearly with the light frequency? (2) is the proportionality factor h the same for all substances? The quantitative test of the theory required some ten years. There were experimental difficulties connected with preparing metal surfaces which were free of impurities (for example, a layer of oxidized metal). It was not until 1916 that it was established that there is indeed a straight-line relationship between the frequency of the light and the maximum kinetic energy of the electrons. to the point where the experimental points on the graph fit a straight line obviously better than any other line. (See the figure on the next page.) Having achieved that degree of accuracy, Millikan could then show that the straight lines obtained for different metals all had the same slope, even though the threshold frequencies were different. The value of h could be obtained from Millikan's measurements; it agreed very well with a value obtained by means of another, independent method. So Einstein's theory was verified quantitatively.

See "Space Travel: Problems of Physics and Engineering" in Project Physics Reader 5.

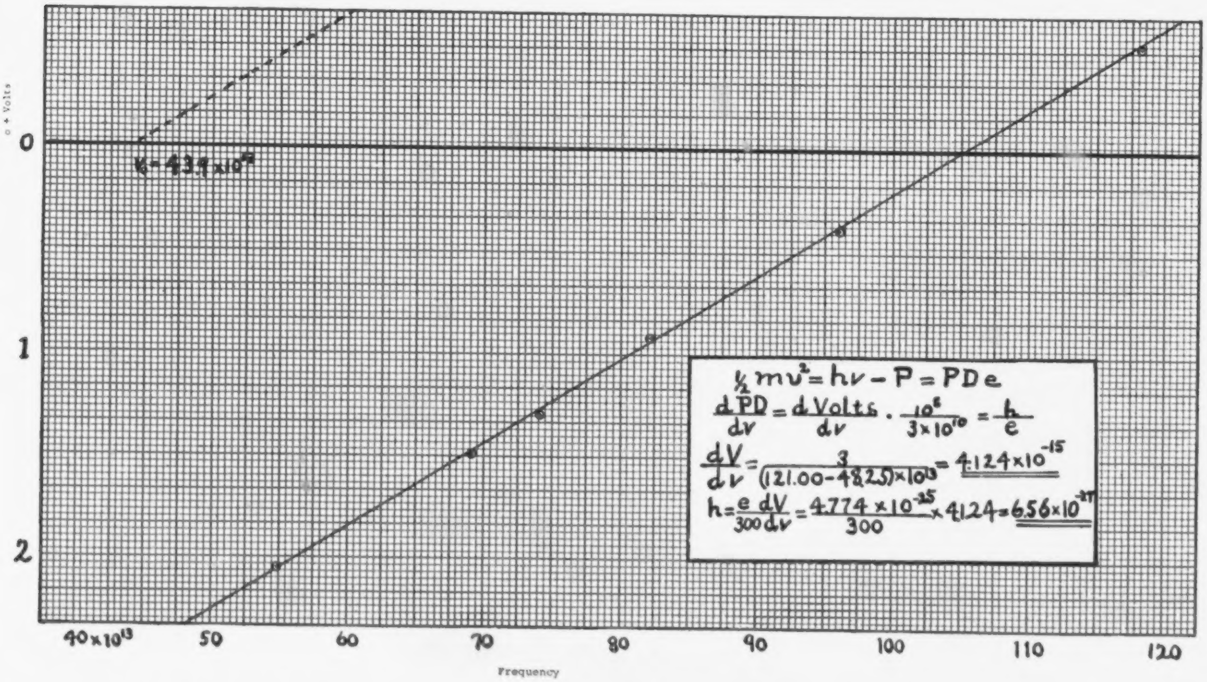
The equation $KE_{max} = hf-W$ led to two Nobel prizes: one to Einstein, who derived it theoretically, and one to Millikan, who verified it experimentally.

SG 18.12

Historically, the first suggestion of a quantum aspect of electromagnetic radiation came from studies of the heat radiated by solids rather than from the photoelectric effect. The concept of quanta of energy hf was introduced by Max Planck, a German physicist, in 1900, five years before Einstein's theory. The constant h is known as Planck's constant. Planck was trying to account for the way in which

Robert Andrews Millikan (1868-1953), an American physicist, attended Oberlin College, where his interest in physics was only mild. After his graduation he became more interested in physics, taught at Oberlin while taking his master's degree, and then obtained his doctor's degree from Columbia University in 1895. After post-doctoral work in Germany he went to the University of Chicago, where he became a professor of physics in 1910. His work on the determination of the electronic charge took from 1906 to 1913. He was awarded the Nobel Prize in physics in 1923 for this research and for the very careful experiments which resulted in the verification of the Einstein photoelectric equation. In 1921, Millikan moved to the California Institute of Technology, eventually becoming its president.

Some of Millikan's data, which verified Einstein's photoelectric equation, are plotted below. The straight-line relationship between frequency and potential is evident and the calculated value of h (in the inset) differs from the best modern values by only one percent. To obtain his data Millikan designed an apparatus in which the metal photoelectric surface was cut clean while in a vacuum. A knife inside the evacuated volume was manipulated by an electromagnet outside the vacuum to make the cuts. This rather intricate arrangement was required to achieve an uncontaminated metal surface.



(Millikan's own symbols are shown here; they are different from the ones used in the text. Can you figure out the meaning of his symbols and units from the information given?)

Millikan had tailing-off problems in finding the stopping voltage, just as students will. One of his solutions was to not extend his curves below a certain minimum current value.

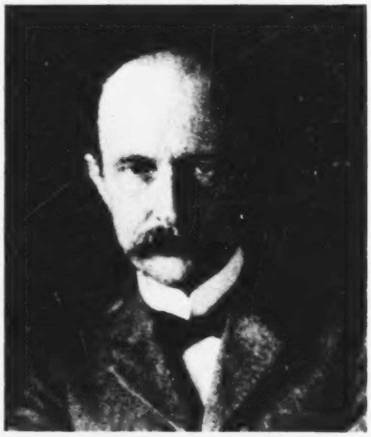
At first sight it is surprizing that a constant so small as h can be detected experimentally and can be a significant factor in most physical and chemical processes. In this connection it must be remembered that h usually occurs in the product hf. The frequency f can be very large.

18.5

the heat energy radiated by a hot body depends on the frequency of the radiation. Classical physics (nineteenth-century thermodynamics and electromagnetism) could not account for the experimental facts. Planck found that the facts could only be interpreted in terms of quanta. Einstein's theory of the photoelectric effect was actually an extension and application of Planck's quantum theory of thermal radiation. Both the experiments on thermal radiation and the theory are much more difficult to describe than are the experiments and the theory of the photoelectric effect. That is why we have chosen to introduce the new (and difficult) concept of quanta of energy by means of the photoelectric effect.

Planck's application of his theory to the experimental data available in 1900 yielded a value of his constant h. The value of h obtained by Millikan in his experiments agreed very well with Planck's value and had greater precision. Additional, independent methods of determining Planck's constant have been devised: the values obtained with all different methods are in excellent agreement.

The photoelectric effect presented physicists with a difficult problem. According to the classical wave theory, light consists of electromagnetic waves extending continuously throughout space. This theory was highly successful in explaining optical phenomena (reflection, refraction, polarization, interference) but could not account for the photoelectric effect. Einstein's theory, in which the existence of discrete bundles of light energy was postulated, accounted for the photoelectric effect; it could not account for the other properties of light. The result was that there were two models whose basic concepts were mutually contradictory. Each model had its successes and failures. The problem was: what, if anything, could be done about the contradictions between the two models? We shall see later that the problem and its treatment have a central position in modern physics.



Max Planck (1858-1947), a German physicist, was the originator of the quantum theory, one of the two great revolutionary physical theories of the 20th century. (The other is Einstein's relativity theory.) Planck won the Nobel Prize in 1918 for his quantum theory. He tried hard to make this theory fit in with the classical physics of Newton and Maxwell, but never succeeded. Einstein extended the idea of quanta much further than Planck himself did.

Surprisingly, Planck was skeptical of Einstein's photoelectric theory when it was first introduced and once said, "If in some of his speculations—as for example in his hypothesis of the light quanta—he was overshooting the target, this should hardly be counted against him. Without taking certain risks one would not be able to advance even in the most exact of the sciences." In spite of this early disagreement Planck and Einstein were friends and had the greatest respect for each other's scientific achievements.

Q7 Einstein's idea of a quantum of light had a definite relation to the wave model of light. What was it?

Q8 Why doesn't the electron have as much energy as the quantum of light which ejects it?

Q9 What does a "stopping voltage" of 2.0 volts indicate?

18.6 X rays. In 1895, another discovery was made which, like the photoelectric effect, did not fit in with accepted ideas about electromagnetic waves and needed quanta for its ex-

Summary 18.6

1. The study of cathode rays led to the discovery of x rays.

2. These were found to be electromagnetic waves with a wavelength much shorter than that of visible light.

3. Due to their short wavelength and penetrating power, x rays have many applications.

A footnote from Röntgen's first report: "For brevity's sake I shall use the expression "rays"; and to distinguish them from others of this name I shall call them x rays." He had been referring to the "active agent."

Wilhelm Konrad Röntgen
(1845-1923)

planation. The discovery was that of x rays by the German physicist, Wilhelm Röntgen. The original discovery, its consequences for atomic physics, and the uses of x rays are all dramatic and important. We shall, therefore, discuss x rays in some detail.

On November 8, 1895, Röntgen was experimenting with cathode rays, as were many physicists all over the world. According to a biographer,

> He had covered the all-glass pear-shaped tube with pieces of black cardboard, and had darkened the room in order to test the opacity of the black paper cover. Suddenly, about a yard from the tube, he saw a weak light that shimmered on a little bench he knew was nearby. Highly excited, Röntgen lit a match and, to his great surprise, discovered that the source of the mysterious light was a little barium platinocyanide screen lying on the bench.

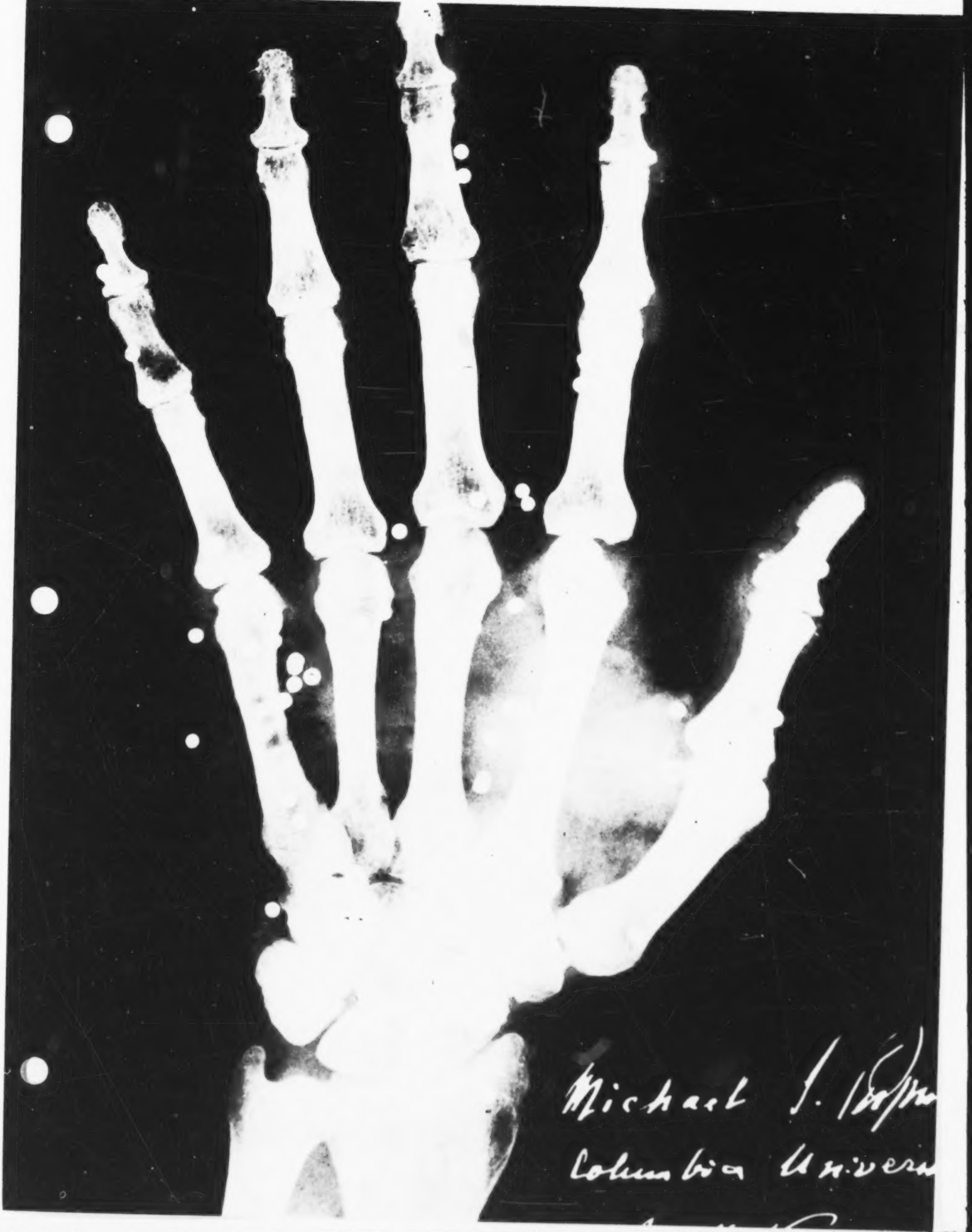
Barium platinocyanide, a mineral, is one of the many chemicals known to fluoresce, that is, to emit visible light when illuminated with ultraviolet light. No source of ultraviolet light was present in Röntgen's experiment. Cathode rays had not been observed to travel more than a few centimeters in air. Hence, neither ultraviolet light nor the cathode rays themselves could have caused the fluorescence. Röntgen, therefore, deduced that the fluorescence he had observed was due to rays of a new kind, which he named x rays, that is rays of an unknown nature. During the next seven weeks he made a series of experiments to determine the properties of this new radiation. He reported his results on Dec. 28 1895 to the Würtzberg Physical Medical Society in a paper whose title, translated, is "On a New Kind of Rays."

The discovery of x rays was narrowly missed by several physicists. Hertz and Lenard (another well-known German physicist) failed to distinguish the cathode rays—perhaps because they didn't happen to have a piece of paper covered with barium platinocyanide lying around to set them on the track. An English physicist, Frederick Smith, found that photographic plates kept in a box near a cathode-ray tube were liable to be fogged—he told his assistant to keep them in another place!

Röntgen's paper described nearly all of the properties of x rays that are known even now. It included an account of the method of production of the rays and proof that they originated in the glass wall of the tube where the cathode rays strike. Röntgen showed that the rays travel in straight lines from their place of origin and that they darken a photographic plate. He reported in detail the ability of the x rays to penetrate various substances—paper, wood, aluminum, platinum and lead. Their penetrating power was greater through light materials (paper, wood, flesh) than through dense materials (platinum, lead, bone). He described photographs showing "the shadows of bones of the hand, of a set of weights in a small box, and of a piece of metal whose inhomogeneity becomes apparent with x rays.' He gave a clear description of the shadows cast by the bones of the hand on the fluorescent screen. Röntgen also reported that the x rays

See TG, p.47 for brief note on Röntgen.

Opposite: One of the earliest x-ray photographs made in the United States. It was made by Michael Pupin of Columbia University in 1896. The man x rayed had been hit by a shotgun blast.



Michael I. Pup

Columbia Univers

were not deflected by a magnetic field, and showed no reflection, refraction or interference effects in ordinary optical apparatus.

One of the most important properties of x rays was discovered by J. J. Thomson a month or two after the rays themselves had become known. He found that when the rays pass through a gas they make it a conductor of electricity. He attributed this effect to "a kind of electrolysis, the molecule being split up, or nearly split up by the Röntgen rays." The x rays, in passing through the gas, knock electrons loose from some of the atoms of the gas. The atoms that lose these electrons become positively charged. They are called ions because they resemble the positive ions in electrolysis, and the gas is said to be ionized.

X rays were often referred to as Röntgen rays after their discoverer.

Röntgen and Thomson found, independently, that the ionization of air produced by x rays discharges electrified bodies. The rate of discharge was shown to depend on the intensity of the rays. This property was therefore used as a quantitative means of measuring the intensity of an x-ray beam. As a result, careful quantitative measurements of the properties and effects of x rays could be made.

One of the problems that aroused interest during the years following the discovery of x rays was that of the nature of the rays. They did not act like charged particles—electrons for example—because they were not deflected by a magnetic field (or by an electric field). They therefore had to be either neutral particles or electromagnetic waves. It was difficult to choose between these two possibilities. On the one hand, no neutral particles of atomic size (or smaller) were known which had the penetrating power of x rays. The existence of neutral particles with high penetrating power would be extremely hard to prove in any case, because there was no way of getting at them. On the other hand, if the x rays were electromagnetic waves, they would have to have extremely short wavelengths: only in this case, according to theory, could they have high penetrating power and show no refraction or interference effects with optical apparatus.

Such a particle—the neutron—was discovered in 1932. You will see in Chapter 23 (Unit 6) how hard it was to identify. But the neutron has nothing to do with x rays.

In 1906 x rays were established as transverse waves by Charles Glover Barkla (British, 1877-1944) by means of a polarization experiment.

The spacing between atoms in a crystal is very small. It was thought, therefore, that if x rays were waves, they would show diffraction effects when transmitted through crystals. In 1912, experiments on the diffraction of x rays by crystals showed that x rays do, indeed, act like electromagnetic radiations of very short wavelength—like ultra ultraviolet light. These experiments are too complicated to

About 1912 Max von Laue showed that x rays could be diffracted, and he thus helped to establish the wave nature of this radiation. 18.6

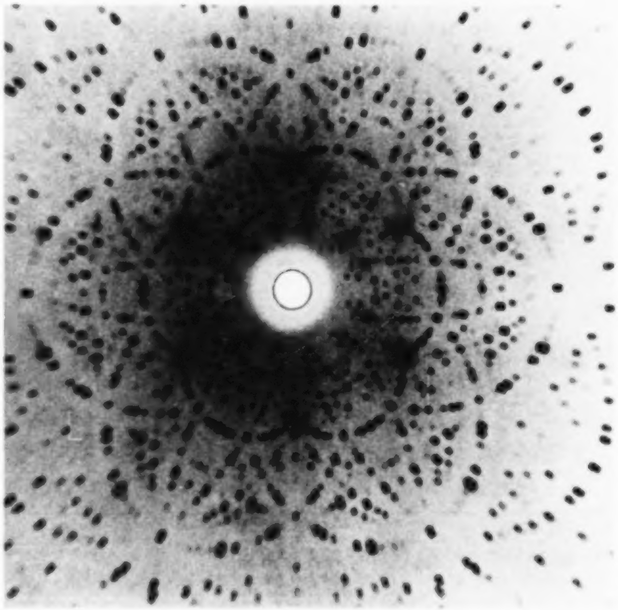
discuss here but they were convincing to physicists, and the problem of the nature of x rays seemed to be solved.

X rays were also found to have quantum properties. They cause the emission of electrons from metals. These electrons have greater kinetic energies than those produced by ultraviolet light. The ionization of gases by x rays is also an example of the photoelectric effect; in this case the electrons are freed from molecules. Thus, x rays also require quantum theory for the explanation of their behavior. The problem of the apparent wave and particle properties of light was aggravated by the discovery that x rays also showed wave and particle properties.

Röntgen's discovery excited intense interest throughout the entire scientific world. His experiments were repeated, and extended, in many laboratories in both Europe and America. The scientific journals, during the year 1896, were filled with letters and articles describing new experiments or confirming the results of earlier experiments. This widespread experimentation was made possible by the fact that, during the years before Röntgen's discovery, the passage of electricity through gases had been a popular topic for study by physicists. Hence many physics laboratories had cathode-ray tubes and could produce x rays easily. Intense interest in x rays was generated by the spectacular use of these rays in medicine. Within three months of Röntgen's discovery, x rays were being put to practical use in a hospital in Vienna in connection with surgical operations. The use of this new aid to surgery spread rapidly. Since Röntgen's time, x rays have revolutionized certain phases of medical practice, especially the diagnosis of some diseases and the treatment of cancer. In other fields of applied science, both physical and biological, uses have been found for x rays which are nearly as important as their use in medicine. Among these are the study of the crystal structure of materials; "industrial diagnosis," such as the search for possible defects in the materials of engineering; the analysis of such different substances as coal and corn; the study of old paintings; the detection of artificial gems; the study of the structure of rubber; and many others.

The position of the atoms in the crystal can be inferred from a diffraction pattern. See figure below. The pattern depends upon the orientation of the crystal, so when the orientation of the crystal relative to the x-ray beam is changed, the pattern changes.

X-ray diffraction patterns from a metal crystal.



Another example of x-ray diffraction is found on p. 100.

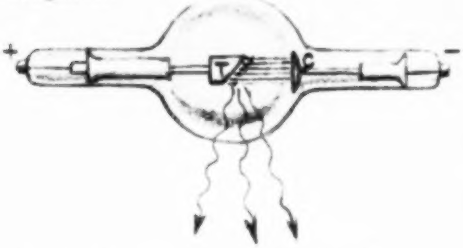
An English physicist, Sir Arthur Schuster, wrote that for some time after the discovery of x rays, his laboratory at Manchester was crowded with medical men bringing patients who were believed to have needles in various parts of their bodies.

"An act to prohibit the insertion of x rays or any device for producing the same into, or their use in connection with, opera glasses or similar aids to vision."

Title of a bill introduced into one of our state Legislatures in 1896.

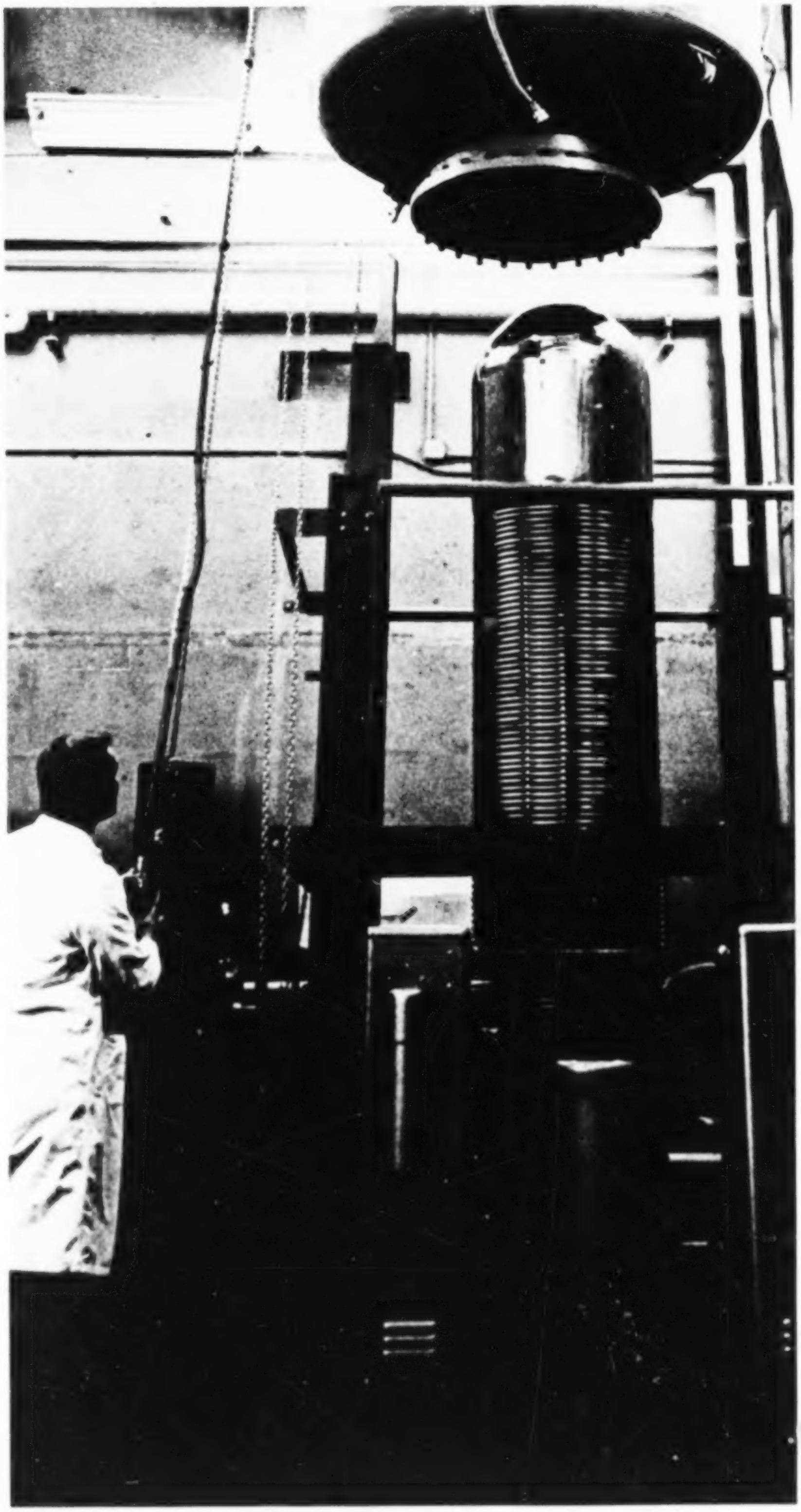
This phenomenon is known as Bremsstrahlung - a German word meaning "braking radiation."

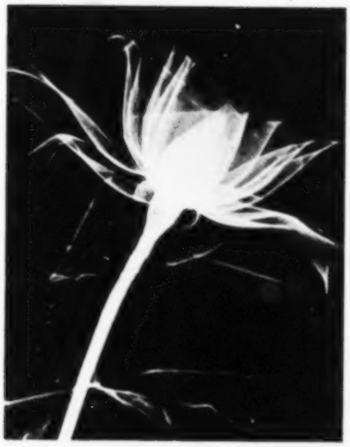
X rays are commonly produced by directing a beam of high energy electrons onto a metal target. As the electrons are deflected and stopped, x rays of various energies are produced. The maximum energy a single x ray can have is the total kinetic energy of an incident electron. So the greater the voltage across which the electron beam is accelerated, the more energetic—and penetrating—are the x rays. One type of x ray tube is shown in the sketch below, where a stream of electrons is emitted from C and accelerated across a high voltage to a tungsten target T.

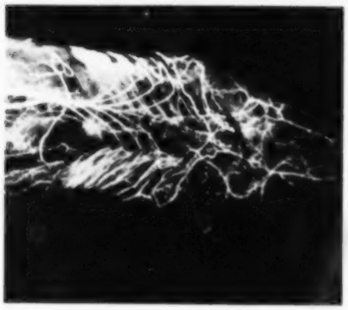


In the photograph at the right is a high voltage machine which is used to produce x rays for research. This van de Graaf type of generator (named after the American physicist who invented it), although not very different in principle from the electrostatic generators of the 1700's, can produce an electric potential difference of 4,000,000 volts.

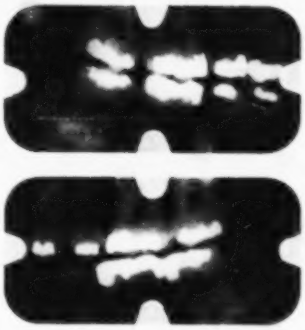
Such a high voltage is possible because of a container, seen in the photograph about to be lowered over the generator, which will be filled with a nonconducting gas under high pressure. (Ordinarily, the strong electric fields around the charged generator would ionize the air and charge would leak off.)

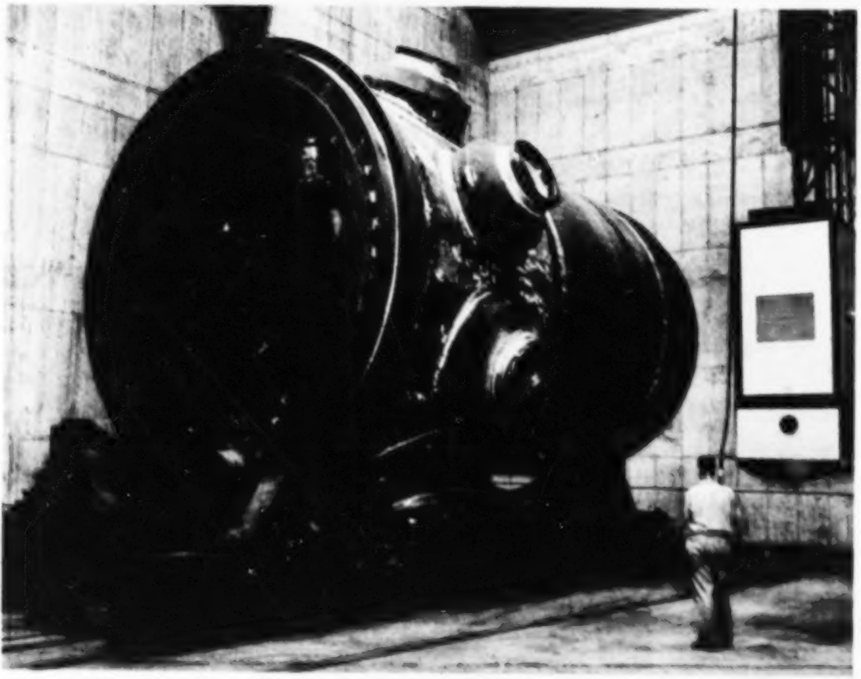






Above left is a rose, photographed with x rays produced by an accelerator-voltage of 30,000 volts. At left is the head of a dogfish shark; its blood vessels have been injected with a fluid that absorbs x rays. Below, x rays are being used to inspect the welds of a 400-ton tank for a nuclear reactor. At the right is the familiar use of x rays in dentistry and the resulting records. Because x rays are injurious to tissues, a great deal of caution is required in using them. For example, the shortest possible pulse is used, lead shielding is provided for the body, and the technician retreats behind a wall of lead and lead glass.





Q10 X rays were the first "ionizing" radiation discovered. What does "ionizing" mean?

Q11 What were three properties of x rays that led to the conclusion that x rays were electromagnetic waves?

Q12 What was the evidence that x rays had a very short wavelength?

Summary 18.7
1. With the electron strongly suspected as a constituent of atoms, Thomson proposed the first model of the atom.

2. Thomson postulated that the electrons were arranged in a balanced pattern inside a sphere of positive charge. The amounts of positive and negative charge were equal so the atom was electrically neutral.

18.7 Electrons, quanta and the atom. By the beginning of the twentieth century enough chemical and physical information was available so that many physicists devised models of atoms. It was known that electrons could be obtained from many different substances and in different ways. But, in whatever way the electrons were obtained, they were always found to have the same properties. This suggested the notion that electrons are constituents of all atoms. But electrons are negatively charged, while samples of an element are ordinarily neutral and the atoms making up such samples are also presumably neutral. Hence the presence of electrons in an atom would require the presence also of an equal amount of positive charge.

The determination of the values of q/m for the electron and for charged hydrogen atoms (ions, in electrolysis experiments) indicated, as mentioned in Sec. 18.2, that hydrogen atoms are nearly two thousand times more massive than electrons. Experiments (which will be discussed in some detail in Chapter 22) showed that electrons constitute only a very small part of the atomic mass in atoms more massive than those of hydrogen. Consequently any model of an atom must take into account the following information: (a) an electrically neutral atom contains equal amounts of positive and negative charge; (b) the negative charge is associated with only a small part of the mass of the atom. Any atomic model should answer two questions: (1) how many electrons are there in an atom, and (2) how are the electrons and the positive charges arranged in an atom?

The approximate size of the atom was known at this time, since good estimates existed for Loschmidt's number (see Unit 3, p. 90).

During the first ten years of the twentieth century several atomic models were proposed, but none was satisfactory. The early models were all based upon classical physics, that is, upon the physics of Newton and Maxwell. No one knew how to invent a model based upon the theory of Planck which incorporated the quantization of energy. There was also need for more experimental knowledge. Nevertheless this state of affairs didn't keep physicists from trying: even a partly wrong model might suggest experiments that might, in turn, provide clues to a better model. Until 1911 the most popular model was one proposed by J. J. Thomson in 1904.

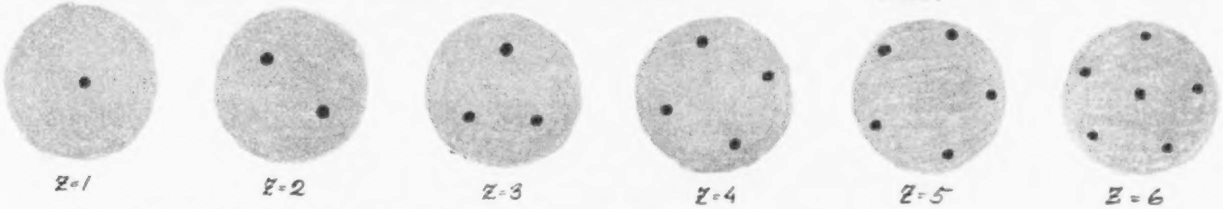
18.7

L47: Thomson model of the atom

Thomson suggested that an atom consisted of a sphere of positive electricity in which was distributed an equal amount of negative charge in the form of electrons. Under this assumption, the atom was like a pudding of positive electricity with the negative electricity scattered in it like raisins. The positive "fluid" was assumed to act on the negative charges, holding them in the atom by means of electric forces only. Thomson did not specify how the positive "fluid" was held together. The radius of the atom was taken to be of the order of 10^{-8} cm on the basis of information from the kinetic theory of gases and other considerations. With this model Thomson was able to calculate certain properties of atoms. For example, he could calculate whether it would be possible for a certain number of electrons to remain in equilibrium, that is, to stay inside the atom without flying apart. Thomson found that certain arrangements of electrons would be stable. Thus, Thomson's model was consistent with the existence of stable atoms. Thomson's theory also suggested that chemical properties might be associated with particular groupings of electrons. A systematic repetition of chemical properties might then occur among groups of elements. But it was not possible to deduce the structure of particular elements and no detailed comparison with the actual periodic table could be made.

See "The 'Thomson' Atom" in Project Physics Reader 5.

Be sure to emphasize that this was only conjecture, and did in fact prove to be wrong.



Some stable arrangements of electrons in Thomson atoms. The atomic number Z is interpreted as the number of electrons.

In Chapter 19 we shall discuss some additional experimental information that provided valuable clues to the structure of atoms. We shall also see how one of the greatest physicists of our time, Niels Bohr, was able to combine the experimental evidence with the concept of quanta into an exciting theory of atomic structure. Although Bohr's theory was eventually replaced, it provided the clues that led to the presently accepted theory of the atom—the quantum mechanical theory.

Q13 Why was most of the mass of an atom believed to be associated with positive electric charge?

18.1 In Thomson's experiment (Fig. 18.4) on the ratio of charge to mass of cathode ray particles, the following might have been typical values for B, V and d: with a magnetic field B alone the deflection of the beam indicated a radius of curvature of the beam within the field of 0.114 meters for $B = 1.0 \times 10^{-3}$ tesla. With the same magnetic field, the addition of an electric field in the same region (V = 200 volts, plate separation d = 0.01 meter) made the beam come straight through.

a) Find the speed of the cathode ray particles in the beam. 2.0×10^{7} m/sec
b) Find q/m for the cathode ray particles. 1.8×10^{11} coul/Kg

The MKSA unit of B is $\frac{N}{\text{amp-m}}$ and is now called the tesla (after the electrical engineer, Nikola Tesla). Measured in this unit the earth's magnetic field is about 0.00005 T and that of a good electromagnet about 1.0t.

formerly Weber/m²

18.2 Given the value for the charge on the electron, show that a current of one ampere is equivalent to the movement of 6.25×10^{18} electrons per second past a given point. derivation

18.3 In the apparatus of Fig. 18.7, an electron is turned back before reaching plate A and eventually arrives at plate C from which it was ejected. It arrives with some kinetic energy. How does this final energy of the electron compare with the energy it had as it left the cathode? Same value

18.4 It is found that at light frequencies below the critical frequency no photoelectrons are emitted. What happens to the light energy? Either absorbed as a whole, or reflected.

Planck's constant has the value $h = 6.6 \times 10^{-34}$ joule-sec.

18.5 For most metals, the work function W is about 10^{-19} joules. To what frequency does this correspond? In what region of the spectrum is this frequency? 1.5×10^{14} cycles/sec Ultra-violet

18.6 What is the energy of a light photon which has a wavelength of 5×10^{-7} m? 5×10^{-8} m? 4×10^{-19} j and 4×10^{-18} j

18.7 The minimum or threshold frequency for emission of photoelectrons for copper is 1.1×10^{15} cycles/sec. When ultraviolet light of frequency 1.5×10^{15} cycles/sec shines on a copper surface, what is the maximum energy of the photoelectrons emitted, in joules? In electron volts? 2.6×10^{-19} j or 1.6 eV

18.8 What is the lowest-frequency light that will cause the emission of photoelectrons from a surface for which the work function is 2.0 eV, that is, a surface such that at least 2.0 eV of energy are needed to eject an electron from it? 4.9×10^{14} cycles/sec

18.9 Monochromatic light of wavelength 5000 Å falls on a metal cathode to produce photoelectrons. The light intensity at the surface of the metal is 10^{2} joules/m² per sec.

a) How many photons fall on 1 m² in one sec? 2.5×10^{20} photons
b) If the diameter of an atom is 1 Å, how many photons fall on one atom in one second on the average? 2.5 photons/sec
c) How often would one photon fall on one atom on the average? 0.4 sec
d) How many photons fall on one atom in 10^{-10} sec on the average? 2.5×10^{-10} photons
e) Suppose the cathode is a square 0.05 m on a side. How many electrons are released per second, assuming every photon releases a photoelectron? How big a current would this be in amperes? 6.3×10^{17} photons/sec 0.1 amp.

18.10 Roughly how many photons of visible light are given off per second by a 1-watt flashlight? (Only about 5 per cent of the electric energy input to a tungsten-filament bulb is given off as visible light.) 1.3×10^{17} photons

Hint: first find the energy, in joules, of an average photon of visible light.

18.11 The highest frequency, f_{max}, of the x rays produced by an x-ray machine is given by the relation

$$hf_{max} = q_e V,$$

where h is Planck's constant and V is the potential difference at which the machine operates. If V is 50,000 volts, what is f_{max}? 1.2×10^{-19} cycles/sec

18.12 The equation giving the maximum energy of the x rays in the preceding problem looks like one of the equations in Einstein's theory of the photoelectric effect. How would you account for this similarity? discussion

18.13 What potential difference must be applied across an x-ray tube for it to emit x rays with a minimum wavelength of 10^{-11} m? What is the energy of these x rays in joules? In electron volts? 1.2×10^{5} volts
1.9×10^{-14} j or 1.2×10^{5} eV

18.14 A glossary is a collection of terms limited to a special field of knowledge. Make a glossary of terms that appeared for the first time in this course in Chapter 18. Make an informative statement about each concept. discussion

18.15 In his Opticks, Newton proposed a set of hypotheses about light which, taken together, constitute a fairly complete model of light. The hypotheses were stated as questions. Three of the hypotheses are given below:

> Are not all hypotheses erroneous, in which light is supposed to consist in pression or motion waves ...? [Quest. 28]
>
> Are not the rays of light very small bodies emitted from shining substances? [Quest. 29]
>
> Are not gross bodies and light convertible into one another, and may not bodies receive much of their activity from the particles of light which enter their composition? [Quest. 30]

a) Was Einstein's interpretation of the photoelectric effect anticipated by Newton? How are the models similar? How different? discussion

b) Why would Newton's model be insufficient to explain the photoelectric effect? What predictions can we make with Einstein's model that we can't with Newton's? discussion

Some suggestion of Millikan's character may be inferred from a standard joke among physicists to the effect that the name Millikan should be interpreted as a thousandth of a "kan" (as in millimeter), where "one kan" is a unit of scientific ability (as in "I can").

Chapter 19 The Rutherford-Bohr Model of the Atom

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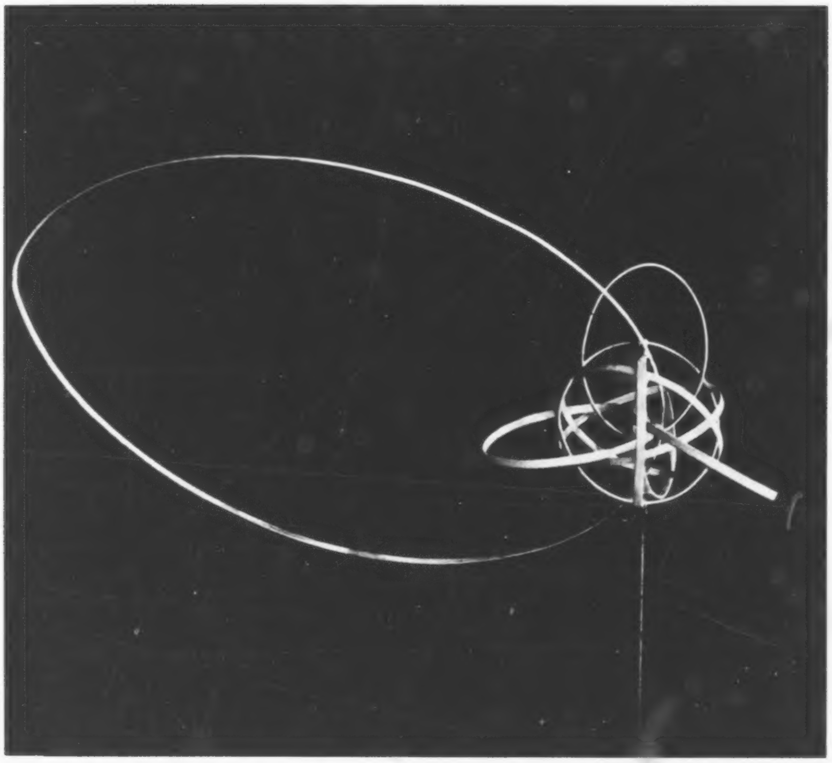


Sculpture representing the Bohr model of a sodium atom.

19.1 Spectra of gases. One of the first real clues to our understanding of atomic structure was provided by the study of the emission and absorption of light by samples of the elements. This study, carried on for many years, resulted in a clear statement of certain basic questions that had to be answered by any theory of atomic structure, that is, by any atomic model. The results of this study are so important to our story that we shall review the history of their development in some detail.

Summary 19.1
1. Elements can be uniquely identified by their spectra. These spectra can be described by relatively simple, empirical formulas. The first such formula, the Balmer formula, describes a spectral series for hydrogen which occurs in the visible region. Balmer's formula led to the suggestion that spectral series should exist in both the infrared and ultraviolet.

It had long been known that light is emitted by gases or vapors when they are excited in any one of several ways: by heating the gas to a high temperature, as when some volatile substance is put into a flame; by an electric discharge, as when the gas is between the terminals of an electric arc or spark; by a continuous electric current in a gas at low pressure, as in the familiar "neon sign."

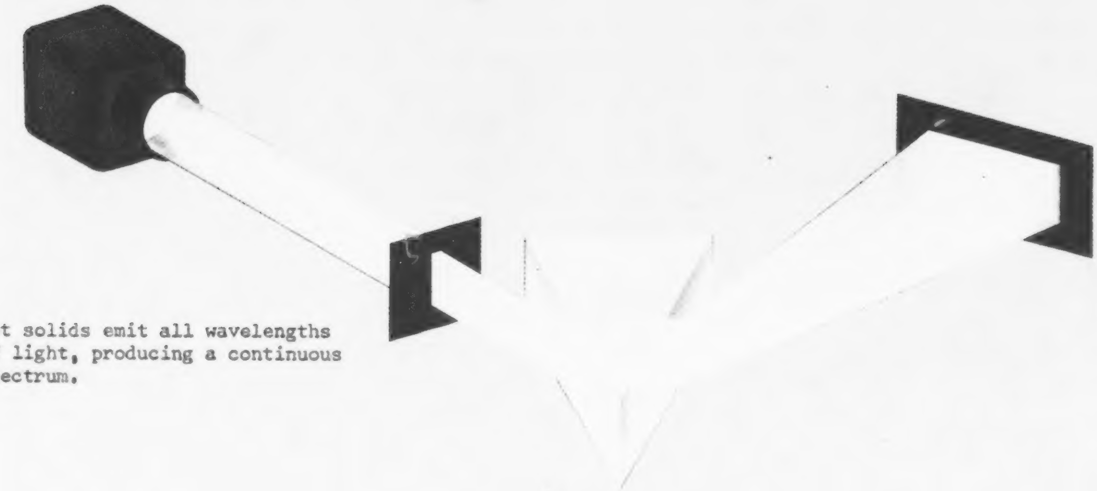
The pioneer experiments on light emitted by various excited gases were made in 1752 by the Scottish physicist Thomas Melvill. He put one substance after another in a flame; and "having placed a pasteboard with a circular hole in it between my eye and the flame..., I examined the constitution of these different lights with a prism." Melvill found the spectrum of light from a hot gas to be different from the continuum of rainbow colors in the spectrum of a glowing solid or liquid. Melvill's spectrum consisted, not of an unbroken stretch of color continuously graded from violet to red, but of individual circular spots, each having the color of that part of the spectrum in which it was located, and with dark gaps (missing colors) between the spots. Later, when more general use was made of a narrow slit through which to pass the light, the spectrum of a gas was seen as a set of lines (Fig. 19.1); the lines are colored images of the slit. Thus the spectrum of light from a gas came to be called a line emission spectrum. From our general theory of light and of the separation of light into its component colors by a prism, we may infer that light from a gas is a mixture of only a few definite colors or narrow wavelength regions of light.

E43*: Spectroscopy

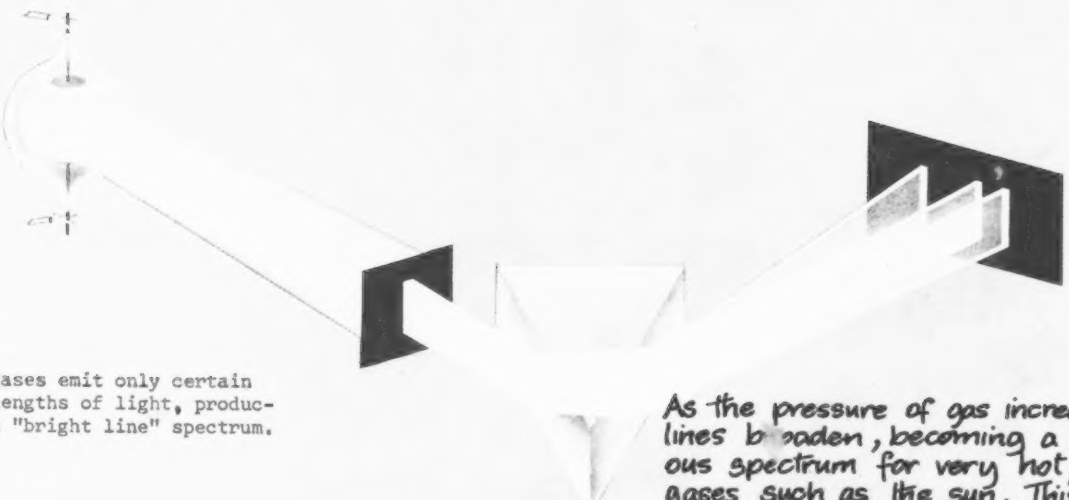
D56: Blackbody radiation

Diffraction gratings are now generally used rather than prisms, but the idea is more easily explained with prisms.

Melvill also noted that the colors and locations of the bright spots were different when different substances were put in the flame. For example, with ordinary table salt in the flame, the predominant color was "bright yellow" (now known to be characteristic of the element sodium). In fact, the line emission spectrum is markedly different for each chemically different gas. Each chemical element has its own characteristic set of wavelengths (Fig. 19.1). In

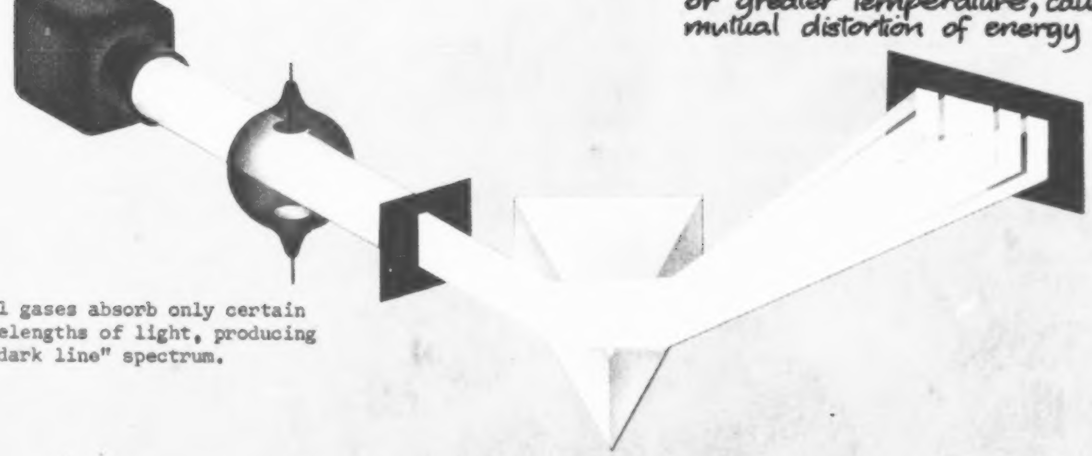


Hot solids emit all wavelengths of light, producing a continuous spectrum.



Hot gases emit only certain wavelengths of light, producing a "bright line" spectrum.

As the pressure of gas increases the lines broaden, becoming a continuous spectrum for very hot, dense gases such as the sun. This point can be explained after Sec. 19.7—the close approach of atoms which accompanies greater density and/or greater temperature, causes mutual distortion of energy levels.



Cool gases absorb only certain wavelengths of light, producing a "dark line" spectrum.

looking at a gaseous source without the aid of a prism or a grating, the eye synthesizes the separate colors and perceives the mixture as reddish for glowing neon, pale blue for nitrogen, yellow for sodium vapor, and so on.

Some gases have relatively simple spectra. Thus sodium vapor shows two bright yellow lines in the visible part of the spectrum. Modern measurements give 5889.953 Å and 5895.923 Å for their wavelengths. Only a good spectrometer can separate them clearly, and we usually speak of them as a sodium "doublet" at about 5890 Å. Some gases or vapors, on the other hand, have exceedingly complex spectra. Iron vapor, for example, has some 6000 bright lines in the visible range alone.

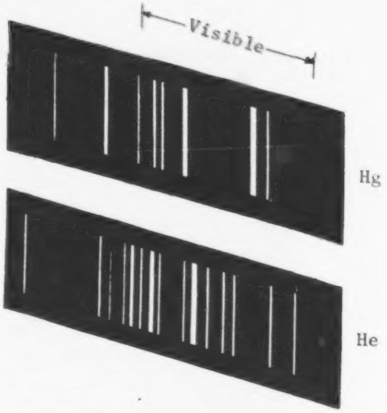


Fig. 19.1 Parts of the line emission spectra of mercury (Hg) and helium (He), redrawn from photographic records.

Some spectra are very complicated. That of vaporized iron contains many thousands of lines.

In 1823 the British astronomer John Herschel suggested that each gas could be identified from its unique line spectrum. Here was the beginning of what is known as spectrum analysis. By the early 1860's the physicist Gustav R. Kirchhoff and the chemist Robert W. Bunsen, in Germany, had jointly discovered two new elements (rubidium and cesium) by noting previously unreported emission lines in the spectrum of the vapor of a mineral water. This was the first of a series of such discoveries; it started the development of a technique making possible the speedy chemical analysis of small samples by spectroscopy.

In 1802 the English scientist William Wollaston saw in the spectrum of sunlight something that had been overlooked before. Wollaston noticed a set of seven sharp, irregularly spaced dark lines across the continuous solar spectrum. He did not understand why they were there, and did not carry the investigation further. A dozen years later, Fraunhofer, the inventor of the grating spectrometer, used better instruments and detected many hundred such dark lines. To the most prominent dark lines, Fraunhofer assigned the letters A, B, C...

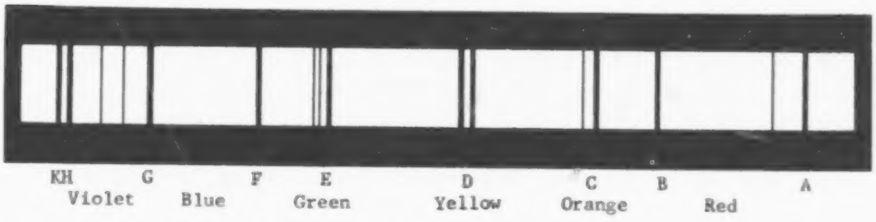


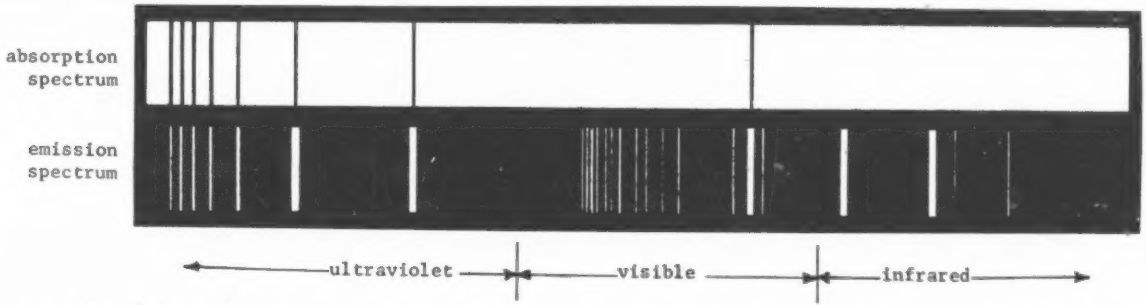
Fig. 19.2 The Fraunhofer dark lines in the visible part of the solar spectrum; only a few of the most prominent lines are represented.

In the spectra of several other bright stars, he found similar dark lines, many of them, although not all, being in the same positions as those in the solar spectrum.

19.1

The key observations toward a better understanding of *both the dark-line and the bright-line spectra of gases* were made by Kirchhoff in 1859. By that time it was known that the two prominent yellow lines in the emission spectrum of heated sodium vapor had the same wavelengths as two prominent dark lines in the solar spectrum to which Fraunhofer had assigned the letter D. It was also known that the light emitted by a glowing solid forms a perfectly continuous spectrum that shows no dark lines. Kirchhoff now demonstrated that if the light from a glowing solid, as on page 66, is allowed first to pass through sodium vapor having a temperature lower than that of the solid emitter and is then dispersed by a prism, the spectrum exhibits two prominent dark lines at the same place in the spectrum as the D-lines of the sun's spectrum. When this experiment was repeated with other gases placed between the glowing solid and the prism, each was found to produce its own characteristic set of dark lines. Evidently each gas in some way absorbs light of certain wavelengths from the passing "white" light; hence such a pattern of dark lines is called a line absorption spectrum, to differentiate it from the bright-line emission spectrum which the same gas would send out at a higher temperature. Most interesting of all, Kirchhoff showed that the wavelength corresponding to each absorption line is equal to the wavelength of a bright line in the emission spectrum of the same gas. The conclusion is that a gas can absorb only light of those wavelengths which, when excited, it can emit (Fig. 19.4). But not every emission line is represented in the absorption spectrum.•

Fig. 19.4 Comparison of the line absorption and emission spectra of sodium vapor.



SG 19.1

Q1 What can you infer about light which gives a bright line spectrum?

Q2 How can such light be produced?

Q3 What can you infer about light which gives a dark line spectrum?

Q4 How can such light be produced?

• For example, it is only the Lyman (ultraviolet) series in hydrogen which appears in the absorption spectrum of cool hydrogen. These are transitions out of the ground state. Other series are exceedingly unlikely because of the low probability of hydrogen atoms being in an excited state. At higher temperatures, the atoms are more likely to be in excited states and lines from other series begin to appear in the hydrogen spectrum.

Summary 19.2
1. Balmer thought up a mathematical formula that fitted the visible spectral lines of hydrogen.
2. The form of the formula suggested to Balmer that there might be other series.

19.2 Regularities in the hydrogen spectrum. The spectrum of hydrogen is especially interesting for historical and theoretical reasons. In the visible and near ultraviolet regions, the emission spectrum consists of a series of lines whose positions are indicated in Fig. 19.5. In 1885, a Swiss school teacher, Johann Jakob Balmer, found a simple formula —an empirical relation—which gave the wavelengths of the lines known at the time. The formula is:

$$\lambda = b\left[\frac{n^2}{n^2 - 2^2}\right].$$

Here b is a constant which Balmer determined empirically and found to be equal to 3645.6 Å, and n is a whole number, different for each line. Specifically, n must be 3 for the first (red) line of the hydrogen emission spectrum (named H_α); n = 4 for the second (green) line (H_β); n = 5 for the third (blue) line (H_γ); and n = 6 for the fourth (violet) line (H_δ). Table 19.1 shows the excellent agreement (within 0.02 %) between the values Balmer computed from his empirical formula and previously measured values.

3. Experimental search revealed just such additional series for hydrogen, and similar series for other gases.

Johann Jakob Balmer (1825-1898), a teacher at a girls' school in Switzerland, came to study wavelengths of spectra listed in tables through his interest in mathematical puzzles and numerology.

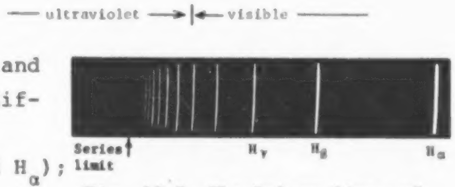


Fig. 19.5 The Balmer lines of hydrogen; redrawn from a photograph made with a film sensitive to ultraviolet light as well as to visible light.

4. The formulas were purely mathematical speculation, without any physical model or reasoning.

In his paper of 1885, Balmer also speculated on the possibility that there might be additional series of hitherto unsuspected lines in the hydrogen spectrum, and that their wavelengths could be found by replacing the 2^2 in the denominator of his equation by other numbers such as 1^2, 3^2, 4^2, and so on. This suggestion, which stimulated many workers to search for such additional spectral series, also turned out to be fruitful. The formula was found to need still another modification (which we shall discuss shortly) before it would correctly describe the new series.

As was the case with Kepler's laws and with periodic law, Balmer's formula gives no hint whatsoever of the underlying mechanism involved. In all these cases a theoretical framework was needed to provide meaning.

To use modern notation, we first rewrite Balmer's formula in a more suggestive form:

$$\frac{1}{\lambda} = R_H\left[\frac{1}{2^2} - \frac{1}{n^2}\right].$$

Name of Line	n	Wavelength λ (Å) From Balmer's formula	By Ångstrom's measurement	Difference
H_α	3	6562.08	6562.10	+0.02
H_β	4	4860.8	4860.74	-0.06
H_γ	5	4340	4340.1	+0.1
H_δ	6	4101.3	4101.2	-0.1

Table 19.1 Data on hydrogen spectrum (as given in Balmer's paper).

19.2

$$R_H = 109{,}677.58\ \text{cm}^{-1}$$

In this equation, which can be derived from the first one, R_H is a constant, equal to 4/b. It is called the Rydberg constant for hydrogen in honor of the Swedish spectroscopist J. R. Rydberg who, following Balmer, made great progress in the search for various spectral series. The lines described by Balmer's formula are said to form a series, called the Balmer series.

If we can now follow Balmer's speculative suggestion of replacing 2^2 by other numbers, we obtain the possibilities:

$$\frac{1}{\lambda} = R_H\left[\frac{1}{1^2} - \frac{1}{n^2}\right]; \qquad \frac{1}{\lambda} = R_H\left[\frac{1}{3^2} - \frac{1}{n^2}\right]; \qquad \frac{1}{\lambda} = R_H\left[\frac{1}{4^2} - \frac{1}{n^2}\right],$$

and so on. All these possible series of lines can be summarized in one formula:

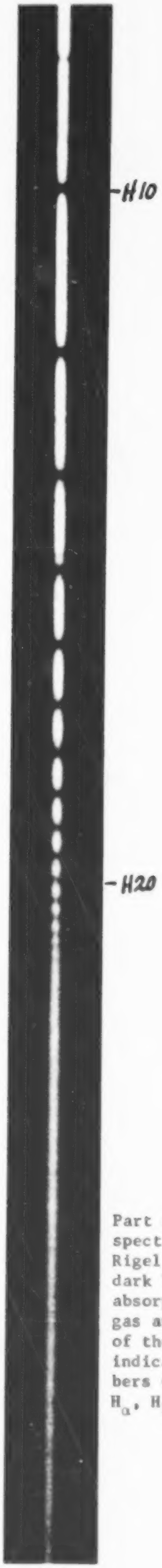
$$\frac{1}{\lambda} = R_H\left[\frac{1}{n_f{}^2} - \frac{1}{n_i{}^2}\right],$$

where n_f is an integer that is fixed for any one series for which wavelengths are to be found (for example, it is 2 for the Balmer series). The letter n_i denotes integers that take on the values $n_f + 1$, $n_f + 2$, $n_f + 3$,... for the successive individual lines in a given series (thus, for the first two lines of the Balmer series, n_i is 3 and 4, respectively). The Rydberg constant R_H should have the same value for all of these hydrogen series.

So far, our discussion has been merely speculation. No series, no single line fitting the formula in the general formula, need exist (except for the Balmer series, where $n_f = 2$). But when we look for these hypothetical lines—we find that they do exist.

In 1908, F. Paschen in Germany found two hydrogen lines in the infrared whose wavelengths were correctly given by setting $n_f = 3$ and $n_i = 4$ and 5 in the general formula, and many other lines in this Paschen series have since been identified. With improvements of experimental apparatus and techniques, new regions of the spectrum could be explored, and then to the Balmer and Paschen series others gradually were added. In Table 19.2 the name of each series is that of its discoverer.

Balmer had also expressed the hope that his formula might indicate a pattern for finding series relationships in the spectra of other gases. This suggestion bore fruit even sooner than the one concerning additional series for hydrogen. Rydberg and others now made good headway in finding



Part of the ultraviolet spectrum of the star Rigel (β Orion). The dark bands are due to absorption by hydrogen gas and match the lines of the Balmer series as indicated by the H numbers (where H_1 would be H_α, H_2 would be H_β etc.).

series formulas for various gases. While Balmer's formula did not serve directly in the description of spectra of gases other than hydrogen, it inspired formulas of similar mathematical form that were useful in expressing order in portions of a good many complex spectra. The Rydberg constant R_H also reappeared in such empirical formulas.

Table 19.2 Series of lines in the hydrogen spectrum.

Name of series	Date of Discovery	Values in Eq. (19.3)	Region of spectrum
Lyman	1906-1914	$n_f = 1$, $n_i = 2, 3, 4, \ldots$	ultraviolet
Balmer	1885	$n_f = 2$, $n_i = 3, 4, 5, \ldots$	ultraviolet-visible
Paschen	1908	$n_f = 3$, $n_i = 4, 5, 6, \ldots$	infrared
Brackett	1922	$n_f = 4$, $n_i = 5, 6, 7, \ldots$	infrared
Pfund	1924	$n_f = 5$, $n_i = 6, 7, 8, \ldots$	infrared

SG 19.2
SG 19.3
SG 19.4

Physicists tried to account for spectra in terms of atomic models. But the great number and variety of spectral lines, even from the simplest atom, hydrogen, made it difficult to do so. Nevertheless, physicists did eventually succeed in understanding the origin of spectra. In this chapter and the next one, we shall get some idea of how this was done.

Q5 What evidence did Balmer have that there were other series of lines in the hydrogen spectrum with terms 3^2, 4^2, etc.?

Q6 Often discoveries result from careful theories (like Newton's) or a good intuitive grasp of phenomena (like Faraday's). What led Balmer to his relation for spectra?

Summary 19.3

1. Rutherford explored the inner structure of atoms by studying the scattering of alpha particles shot at thin metallic foils.

2. From an analysis of the angles through which the alpha particles were scattered, Rutherford deduced that almost all of an atom's mass is concentrated in an extremely small, positively charged nucleus.

19.3 Rutherford's nuclear model of the atom. A new basis for atomic models was provided during the period 1909 to 1911 by Ernest Rutherford (1871-1937), a New Zealander who had already shown a rare ability as an experimentalist at McGill University, Montreal, Canada. He had been invited in 1907 to Manchester University in England, where he headed a productive research laboratory. Rutherford was specially interested in the rays emitted by radioactive substances, in particular in α (alpha) rays. As we shall see in Chapter 20, α rays consist of positively charged particles. These particles are positively charged helium atoms with masses about 7500 times larger than the electron mass. Some radioactive substances emit α particles at a great enough rate and with enough energy so that the particles can be used as

SG 19.5

F40: The structure of atoms

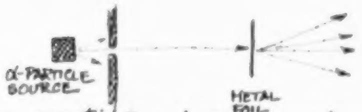
F41: Rutherford atom

projectiles to bombard samples of elements. The experiments that Rutherford and his colleagues did with α particles are an example of a highly important kind of experiment in atomic and nuclear physics—the scattering experiment.

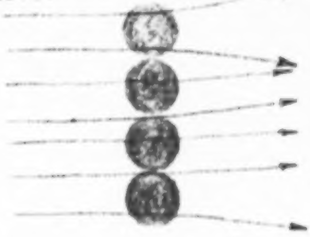
The concept of a planetary atom was not new; it had been suggested by J. Perrin in France in 1901 and H. Nagaoka in Japan in 1904. The proposals for these models always included discussion about the problem of calculating the frequencies of emitted radiation.

See TG, p. 90 for note on Nagaoka's theory of the "Saturnian" atom.

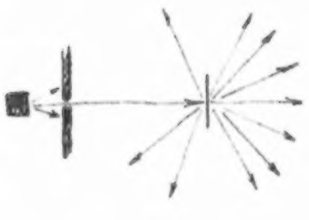
In a scattering experiment, a narrow, parallel beam of projectiles or bullets (α particles, electrons, x rays) is aimed at a target that is usually a very thin foil or film of some material. As the beam strikes the target, some of the projectiles are deflected, or scattered, from their original direction. The scattering is the result of the interaction between the particles or rays in the beam and the atoms of the material. A careful study of the projectiles after they have been scattered can yield information about the projectiles, the atoms, or both—or the interaction between them. Thus if we know the mass, energy and direction of the projectiles, and see what happens to them in a scattering experiment, we can deduce properties of the atoms that scattered the projectiles.



An α particle, whose mass is about seven thousand times greater than the mass of an electron, should be able to sweep electrons right out of its path with practically no deflection.



Rutherford noticed that when a beam of α particles passed through a thin metal foil, the beam spread out. He thought that some of the particles were scattered out of the beam by colliding with atoms in the foil. The scattering of α particles can be described in terms of the electrostatic forces between the positively charged α particles and the charges that make up atoms. Since atoms contain both positive and negative charges, an α particle is subjected to both repulsive and attractive forces as it passes through matter. The magnitude and direction of these forces depend on how near the particle happens to approach to the centers of the atoms past which it moves. When a particular atomic model is postulated, the extent of the scattering can be calculated quantitatively and compared with experiment. In the case of the Thomson atom, calculation showed that the probability that an α particle would be scattered through an angle of more than a few degrees is negligibly small.



One of Rutherford's assistants, H. Geiger, found that the number of particles scattered through large angles, 10° or more, was much greater than the number predicted on the basis of the Thomson model. In fact, one out of about every 8000 α particles was scattered through an angle greater than 90°. This result meant that a significant number of α particles bounced back from the foil. This result was unexpected. Some years later, Rutherford wrote:

The positive blob will exert on an α particle outside of it the same force as it would if concentrated in a point at the center of the blob. The r in $\frac{1}{r^2}$ is the distance from the center, and the repulsive force goes up very quickly as the α approaches the center of the blob — as long as it stays outside the blob. (Once inside the blob, the repulsive force starts to decrease.) In Thomson's big blob, the α would have to be inside, and therefore subject to weak forces. The very large repulsive forces inferred from scattering data would require that the α could get very close to the center of the + blob without getting inside it!

Rutherford is often thought of as "The Newton of the Atom." He liked this. He was once asked by one of his colleagues, rather enviously: "You are always riding the wave, aren't you, Rutherford?" "Why not," said Rutherford, "I made it, didn't I?" and added soberly, "At least to some extent." He once wrote, "I know of no more enthralling adventure than this voyage of discovery into the almost unexplored world of the atomic nucleus."

Ernest Rutherford was born, grew up, and received most of his education in New Zealand. At age 24 he went to Cambridge, England to work at the Cavendish Laboratory under J.J. Thomson. From there he went to McGill University in Canada, then home to be married and back to England again, now to Manchester University. At these universities, and later at the Cavendish Laboratory where he succeeded J.J. Thomson as director, Rutherford performed important experiments on radioactivity, the nuclear nature of the atom, and the structure of the nucleus. Rutherford introduced the names "alpha," "beta" and "gamma" rays, "protons," and "half-life." For his scientific work, Rutherford was knighted and received a Nobel Prize.

In the photograph above, Rutherford holds the apparatus in which he arranged for α particles to bombard nitrogen nuclei—not to study scattering, but to detect actual disintegration of the nitrogen nuclei! (See Sect. 23.3 in Unit 6 Text.)

T38: Alpha scattering

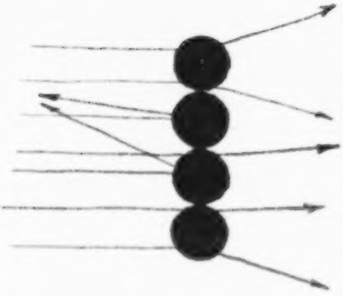
L48: Rutherford scattering

19.3

...I had observed the scattering of α-particles, and Dr. Geiger in my laboratory had examined it in detail. He found, in thin pieces of heavy metal, that the scattering was usually small, of the order of one degree. One day Geiger came to me and said, "Don't you think that young Marsden, whom I am training in radioactive methods, ought to begin a small research?" Now I had thought that, too, so I said, "Why not let him see if any α-particles can be scattered through a large angle?" I may tell you in confidence that I did not believe that they would be, since we knew that the α-particle was a very fast, massive particle, with a great deal of [kinetic] energy, and you could show that if the scattering was due to the accumulated effect of a number of small scatterings, the chance of an α-particle's being scattered backward was very small. Then I remember two or three days later Geiger coming to me in great excitement and saying, 'We have been able to get some of the α-particles coming backward..." It was quite the most incredible event that has ever happened to me in my life. It was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you. On consideration, I realized that this scattering backward must be the result of a single collision, and when I made calculations I saw that it was impossible to get anything of that order of magnitude unless you took a system in which the greater part of the mass of the atom was concentrated in a minute nucleus. It was then that I had the idea of an atom with a minute massive centre, carrying a charge.

Years later, in 1938, an English physicist, C. G. Darwin, wrote, "I count it as one of the great occurrences of my life that I was actually present half-an-hour after the nucleus was born.

"It was a Sunday supper at his [Rutherford's] Manchester house and I remember his saying to us, I would think before supper itself, that he had been looking into the business of big scattering of α particles, and that there must be enormous forces in the atom to do it. Then he said that it would require a charge of something like a hundred electrons in the case of the gold atom, but that may have been added a few days later."



These experiments and Rutherford's idea marked the origin of the modern concept of the nuclear atom. Let us look at the experiments more closely to see why Rutherford concluded that the atom must have its mass and positive charge concentrated at the center, thus forming a nucleus about which the electrons are clustered.

SG 19.5

SG 19.6

A possible explanation of the observed scattering is that there exist in the foil concentrations of mass and charge—positively charged nuclei—much more dense than Thomson's atoms. An α particle heading directly toward one of them is stopped and turned back, as a ball would bounce back from a rock but not from a cloud of dust particles. Figure 19.6 is based on one of Rutherford's diagrams in his paper of 1911, which may be said to have laid the foundation for the modern theory of atomic structure. It shows two α particles A and A'. The α particle A is heading directly toward a nucleus N. Because of the electrical repulsive force between the two, A is slowed to a stop at some distance r from N, and then moves directly back. A' is another α particle that is not headed directly toward the nucleus N; it swerves away from N along a path which calculation showed must be an hyperbola. The deflection of A' from its original path is indicated by the angle ϕ.

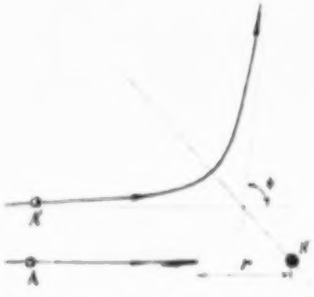


Fig. 19.6 Paths of two α particles A and A' approaching a nucleus N. (Based on Rutherford, Philosophical Magazine, vol. 21 (1911), p. 669.)

See TG, p. 88 for note on Rutherford scattering.

Rutherford calculated theoretically and found experimentally that the number of alpha particles scattered through a given angle ϕ is proportional to $\frac{t\,Q^2}{V_\alpha^4 \sin^4(\phi/2)}$

where:
V_α is initial speed
t foil thickness
Q nuclear charge
ϕ scattering angle

19.3 Rutherford considered the effects of important factors on the α particles—their initial speed v_α, the foil thickness t, and the quantity of charge Q on each nucleus. According to the theory most of the α particles should be scattered through small angles, but a significant number should be scattered through large angles.

Geiger and Marsden undertook tests of these predictions with the apparatus shown schematically in Fig. 19.8. The lead box B contains a radioactive substance (radon) which emits α particles. The particles emerging from the small hole in the box are deflected through various angles ϕ in passing through the thin metal foil F. The number of particles deflected through each angle ϕ is found by letting the particles strike a small zinc sulfide screen S. Each α particle that strikes the screen produces a scintillation (a momentary pinpoint of fluorescence). These scintillations can be observed and counted by looking through the microscope M; S and M can be moved together along the arc of a circle up to ϕ = 150°. In later experiments, the number of α particles at any angle ϕ was counted more conveniently by replacing S and M by a counter (Fig. 19.9) invented by Geiger. The Geiger counter, in its more recent versions, is now a standard laboratory item.

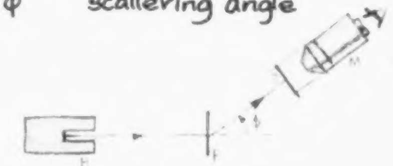


Fig. 19.8 Scintillation method for verifying Rutherford's theoretical predictions for α particle scattering. The whole apparatus is placed in an evacuated chamber so that the α particles will not be slowed down by collisions with air molecules.

Geiger and Marsden found that the number of α particles counted depended on the scattering angle, the speed of the particles, and on the thickness of the foil of scattering material in just the ways that Rutherford had predicted. SG 19.2

Q7 Why should α particles be scattered by atoms?

Q8 What was the basic difference between the Rutherford and the Thomson models of the atom?

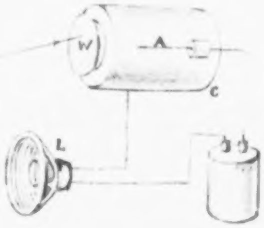


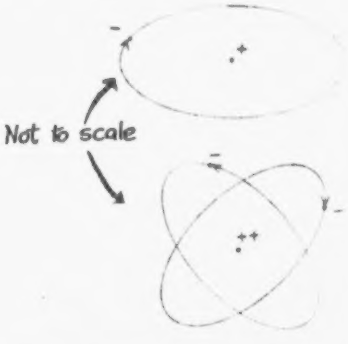
Fig. 19.9 A Geiger counter (1928). It consists of a metal cylinder C containing a gas and a thin axial wire A that is insulated from the cylinder. A potential difference slightly less than that needed to produce a discharge through the gas is maintained between the wire (anode A) and cylinder (cathode C). When an α particle enters through the thin mica window W, it frees a few electrons from the gas molecules, leaving the latter positively charged. The electrons are accelerated toward the anode, freeing more electrons along the way by collisions with gas molecules. The avalanche of electrons constitutes a sudden surge of current which may be amplified to produce a click in the loudspeaker (L).

19.4 Nuclear charge and size. At the time Rutherford made his predictions about the effect of the speed of the α particle and the thickness of foil on the angle of scattering, there was no way independently to measure the charge Q on each nucleus. However, some of Rutherford's predictions were confirmed by scattering experiments and, as often happens when part of a theory is confirmed, it is reasonable to proceed temporarily as if the whole of that theory were justified. Thus it was assumed that the scattering of α particles through a given angle is proportional to the square of the nuclear charge. With this relation in mind, Q could be estimated. Experimental data were obtained for the scattering of different elements. Among them were carbon, aluminum and gold. Therefore, on the basis of this assumption the following nuclear charges were obtained: for carbon $6q_e$, for aluminum 13 or $15q_e$ and for gold 78 or $79q_e$. Similar tentative values were found for other elements.

Summary 19.4
1. Several lines of evidence made it possible to estimate the charge on the nucleus and the number of electrons in an atom.

2. The order of elements in the periodic table corresponds to an increasing number of charges in the atom!

3. Calculations from scattering experiments indicated that the nucleus is extremely small—the atom is mostly empty space.



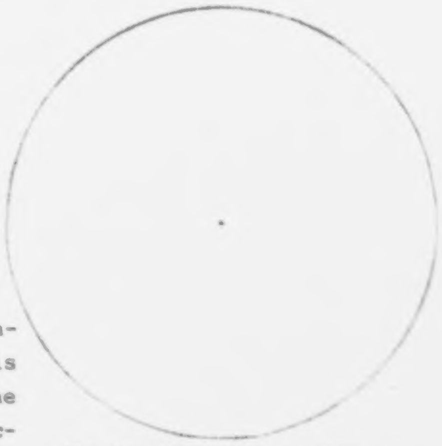
The magnitude of the positive charge of the nucleus was an important piece of information about the atom. If the nucleus has a positive charge of $6\ q_e$, 13 or $14\ q_e$, etc., the number of electrons surrounding the nucleus must be 6 for carbon, 13 or 14 for aluminum, etc., since the atom as a whole is electrically neutral. It was soon noticed that the values found for the nuclear charge were close to the atomic number Z, the place number of the element in the periodic table. The data seemed to indicate that each nucleus has a positive charge Q numerically equal to Zq_e. But the results of experiments on the scattering of α particles were not precise enough to permit this conclusion to be made with certainty.

The suggestion that the number of positive charges on the nucleus and also the number of electrons around the nucleus are equal to the atomic number Z made the picture of the nuclear atom clearer. The hydrogen atom ($Z = 1$) has, in its neutral state, one electron outside the nucleus; a helium atom ($Z = 2$) has in its neutral state two electrons outside the nucleus; a uranium atom ($Z = 92$) has 92 electrons. This simple scheme was made more plausible when additional experiments showed that it was possible to produce singly ionized hydrogen atoms, H^+, and doubly ionized helium atoms, He^{++}, but not H^{++} or He^{+++}, evidently because a hydrogen atom has only one electron to lose, and a helium atom only two. The concept of the nuclear atom provided new insight into the periodic table of the elements: it suggested that the periodic table is really a listing of the elements according to the number of electrons around the nucleus or according to the number of positive units of charge in the nucleus.

Additional evidence for this suggestion was provided by research with x rays during the years 1910 to 1913. It was found that the elements have characteristic x-ray spectra as well as optical spectra. The x-ray spectra show separate lines against a continuous background. A young English physicist, H. G. J. Moseley (1887-1915), found that the frequencies of certain lines in the x-ray spectra of the elements vary in a strikingly simple way with the nuclear charge Z. The combination of the experimental results with the Bohr theory of atomic structure made it possible to assign an accurate value to the nuclear charge of an element. As a result, Moseley established with complete certainty that the place number of an element in the periodic table is the same as the value of the positive charge of the nucleus (in multiples of the unit electric charge) and the same as the number of electrons outside the nucleus. These results made it possible to remove some of the discrepancies in Mendeleev's periodic table and to relate the table in a definite way to the Bohr theory.

H.G.J. Moseley (1887-1915) was a co-worker with Rutherford at Manchester. Bohr characterized him as a man of extraordinary energy and gifts for purposeful experimentation. J.J. Thomson said he made one of the most brilliant discoveries ever made by so young a man. At the start of World War I he volunteered for army service, was sent to the Dardanelles and was killed during the unsuccessful attack at Gallipoli. Rutherford wrote that "it is a national tragedy that our military organization at the start was so inelastic as to be unable, with few exceptions, to utilize the offers of services for scientific men except as combatants on the firing line." In his will Moseley left all his apparatus and private wealth to the Royal Society to promote scientific research.

As an important result of these scattering experiments the size of the nucleus may be estimated. Suppose an α particle is moving directly toward a nucleus (A, Fig. 19.6). Its kinetic energy on approach is transformed into electrical potential energy. It slows down and eventually stops. The distance of closest approach may be computed from the original kinetic energy of the α particle and the charges of α particle and nucleus. It turns out to be approximately 3×10^{-14}m. If the α particle is not to penetrate the nucleus, this distance must be at least as great as the sum of the radii of α particle and nucleus; then the radius of the nucleus could not be larger than about 10^{-14}m, only about 1/1000 of the radius of an atom.• Thus if we consider volumes, which are proportional to cubes of radii, it is clear that the atom is mostly empty space. This must be so to explain the ease with which α particles or electrons penetrate thousands of layers of atoms in metal foils or in gases.



The dot drawn in the middle to represent the nucleus is about 100 times too large. Popular diagrams of atoms often greatly exaggerate the size of the nucleus, to suggest the greater mass.

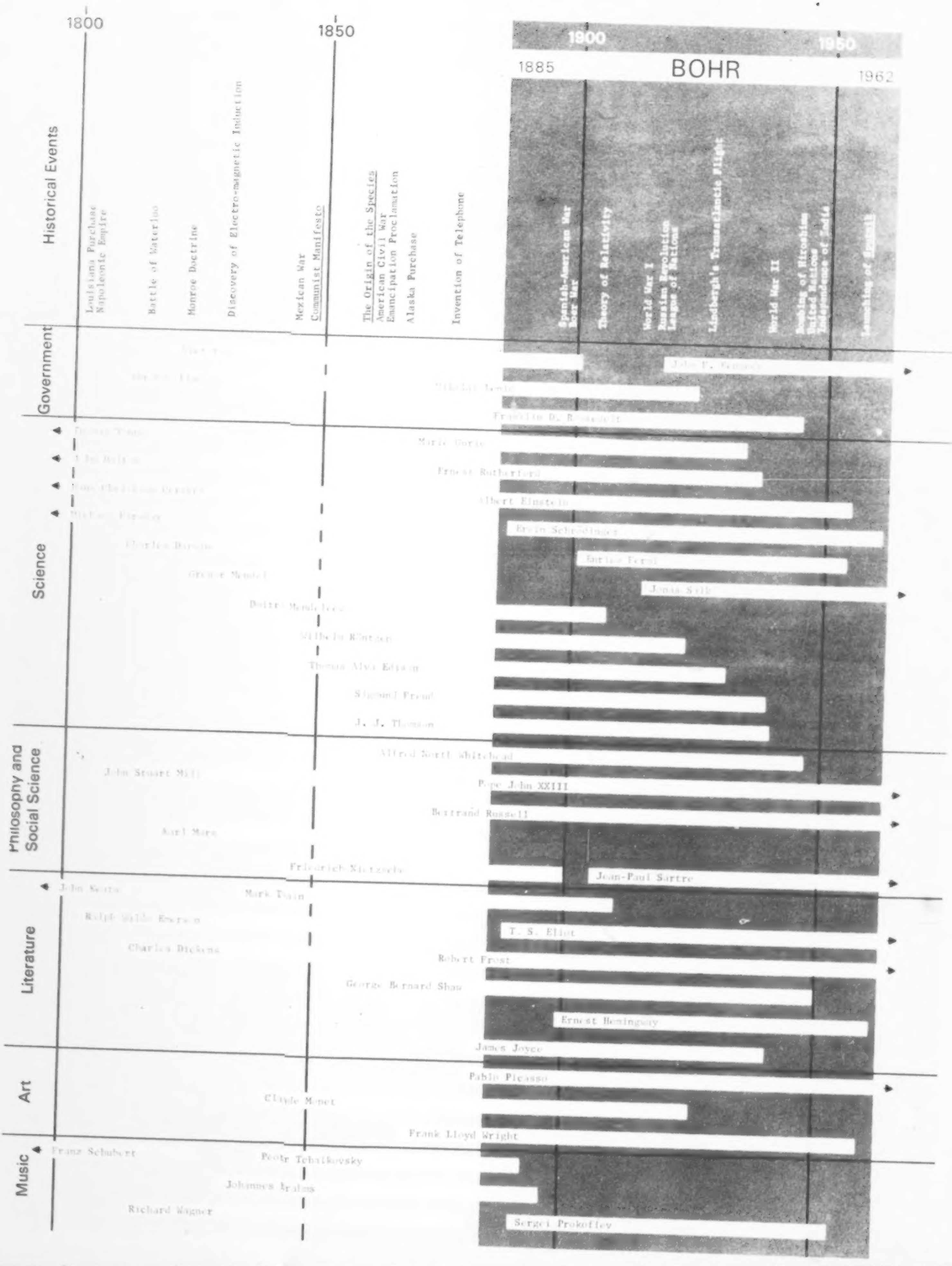
Successful as this model of the nuclear atom was in explaining scattering phenomena, it raised many new questions: What is the arrangement of electrons about the nucleus? What keeps the negative electron from falling into a positive nucleus by electrical attraction? How is the nucleus made up? What keeps it from exploding on account of the repulsion of its positive charges? Rutherford realized the problems raised by these questions and the failure of his model to answer them. Additional assumptions were needed to complete the model—to find answers to the additional questions posed about the details of atomic structure. The remainder of this chapter will deal with the theory proposed by Niels Bohr, a young Danish physicist who joined Rutherford's group just as the nuclear model was being announced.

SG 19.8

Q9 What is the "atomic number" of an element, according to the Rutherford model of the atom?

Q10 What is the greatest electric charge an ion of lithium (the next heaviest element after helium) could have?

•So the nucleus volume is about 10^{-12} times the atomic volume.
If we could squeeze out all the electrons and fill a pint
carton with hydrogen nuclei, the carton would weigh about 10^{8} tons.


1800
1850
1900
1950
1885
BOHR
1962
Historical Events
Louisiana Purchase
Napoleonic Empire
Battle of Waterloo
Monroe Doctrine
Discovery of Electro-magnetic Induction
Mexican War
Communist Manifesto
The Origin of the Species
American Civil War
Emancipation Proclamation
Alaska Purchase
Invention of Telephone
Spanish-American War
Boer War
Theory of Relativity
World War I
Russian Revolution
League of Nations
Lindbergh's Transatlantic Flight
World War II
Bombing of Hiroshima
United Nations
Independence of India
Launching of Sputnik
Government
Victoria
Abraham Lincoln
John F. Kennedy
Nikolai Lenin
Franklin D. Roosevelt
Science
Thomas Young
John Dalton
Hans Christian Oersted
Michael Faraday
Charles Darwin
Gregor Mendel
Dmitri Mendeleev
Wilhelm Röntgen
Thomas Alva Edison
Sigmund Freud
J. J. Thomson
Marie Curie
Ernest Rutherford
Albert Einstein
Erwin Schrödinger
Enrico Fermi
Jonas Salk
Philosophy and Social Science
John Stuart Mill
Karl Marx
Friedrich Nietzsche
Alfred North Whitehead
Pope John XXIII
Bertrand Russell
Jean-Paul Sartre
Literature
John Keats
Ralph Waldo Emerson
Charles Dickens
Mark Twain
T. S. Eliot
Robert Frost
George Bernard Shaw
Ernest Hemingway
James Joyce
Art
Claude Monet
Pablo Picasso
Frank Lloyd Wright
Music
Franz Schubert
Richard Wagner
Johannes Brahms
Peotr Tchaikovsky
Sergei Prokofiev

19.5 The Bohr theory: the postulates. If an atom consists of a positively charged nucleus surrounded by a number of negatively charged electrons, what keeps the electrons from falling into the nucleus—from being pulled in by the Coulomb force of attraction? One possible answer to this question is that an atom may be like a planetary system with the electrons revolving in orbits about the nucleus. Instead of the gravitational force, the Coulomb attractive force between the nucleus and an electron would supply a centripetal force that would tend to keep the electron in an orbit. Although this idea seems to start us on the road to a theory of atomic structure, a serious problem arises concerning the stability of a planetary atom. According to Maxwell's theory of electromagnetism, a charged particle radiates energy when it is accelerated. Now, an electron moving in an orbit around a nucleus is constantly being accelerated by the centripetal force mv^2/r. The electron, therefore, should lose energy by emitting radiation. A detailed analysis of the motion of the electron (which we can't do here because of the mathematical difficulty) shows that the electron should be drawn closer to the nucleus. Within a very short time, the electron should actually be pulled into the nucleus. According to classical physics—mechanics and electromagnetics—a planetary atom would not be stable for more than a small fraction of a second.

Summary 19.5

1. With the nucleus discovered, troublesome questions arose as to the spatial arrangement of the electrons and nucleus within the atom. Classically, the planetary model would be unstable.

2. Bohr postulated a set of stable states (which were specified by allowing certain values of angular momentum). This postulate specified the radii of the possible orbits for the electron, the possible orbital speed, and the possible orbital energies.

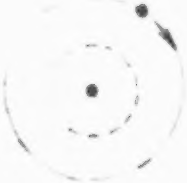
This electron collapse would occur in about 10^{-8} sec.

The idea of a planetary atom was sufficiently attractive that physicists continued to look for a theory that would include a stable planetary structure and predict discrete line spectra for the elements. Bohr succeeded in constructing such a theory in 1913. This theory, although it had to be radically modified later, showed how to attack atomic problems by using quantum theory. In fact, Bohr showed that only by using quantum theory would the problem of atomic structure be attacked with any hope of success. Bohr used the quantum ideas of Planck and Einstein that electromagnetic energy is absorbed or emitted as discrete quanta; and that each quantum has a magnitude equal to Planck's constant h multiplied by the frequency of the radiation.

Bohr introduced two postulates designed to account for the existence of stable electron orbits and of the discrete emission spectra. These postulates may be stated as follows.

(1) An atomic system possesses a number of states in which no emission of radiation takes place, even if the particles (electrons and nucleus) are in motion relative to each other. These states are called stationary states of the atom.

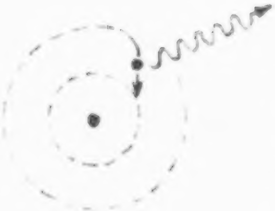
See TG, p.48 for alternate way of stating Bohr's postulates.



(2) Any emission or absorption of radiation, either as visible light or other electromagnetic radiation, will correspond to a transition between two stationary states. The radiation emitted or absorbed in a transition has a frequency f determined by the relation

$$hf = E_i - E_f ,$$

where h is Planck's constant and E_i and E_f are the energies of the atom in the initial and final stationary states, respectively.



These postulates are a combination of some ideas taken over from classical physics together with others in direct contradiction to classical physics. For example, Bohr assumed that when an atom is in one of its stationary states, the motions of the electrons are in accord with the laws of mechanics. A stationary state may be characterized by its energy, or by the orbits of the electrons. Thus, in the simple case of the hydrogen atom, with a single electron revolving about the nucleus, a stationary state corresponds to the electron moving in a particular orbit and having a certain energy. Bohr avoided the difficulty of the electron emitting radiation while moving in its orbit by postulating that it does not emit radiation when it is in a particular orbit. This postulate implies that classical, Maxwellian electromagnetics does not apply to the motion of electrons in atoms. The emission of radiation was to be associated with a jump from a state with one energy (or orbit) to another state with a different energy (or orbit). Bohr did not attempt to explain why the atom should be stable in a given stationary state.

SG 19.9

Bohr's theory includes no attempt to explain the dynamical situation within the atom during the acts of emission or absorption.

The first postulate has in view the general stability of the atom, while the second has (chiefly) in view the existence of spectra with sharp lines. The use of quantum theory enters in the second postulate, and is expressed in the equation $hf = E_i - E_f$. Bohr also used the quantum concept in defining the stationary states of the atom. The states are highly important in atomic theory so we shall look at their definition carefully. For simplicity we consider the hydrogen atom, with a single electron revolving around the nucleus. The positive charge of the nucleus is given by Zq_e with $Z = 1$. We also assume, following Bohr, that the possible orbits of the electron are circles. The condition that the centripetal force is equal to the attractive Coulomb force is:

$$\frac{mv^2}{r} = k\frac{q_e^{\,2}}{r^2} .$$

In this formula, m is the mass of the electron; v is the speed; r is the radius of the circular orbit, that is, the distance of the electron from the nucleus; the nucleus is assumed to be stationary. The symbol k stands for a constant which depends on the units used; q_e is the magnitude of the electronic charge.

The values of r and v which satisfy the centripetal force equation characterize the possible electron orbits. We can write the equation in a slightly different form by multiplying both sides by r^2 and dividing both sides by v; the result is:

$$mvr = k\frac{q_e^2}{v}.$$

The quantity on the left side of this equation, which is the product of the momentum of the electron and the radius of the orbit, can also be used to characterize the orbits. This quantity is often used in problems of circular motion, and it is called the angular momentum:

Angular momentum can be elaborated to taste at this point, but the story line of Unit 5 wouldn't particularly benefit.

angular momentum = mvr.

According to classical mechanics, the radius of the orbit could have any value and the angular momentum could also have any value. But we have seen that under classical mechanics there would be no stable orbits in the hydrogen atom. Since Bohr's first postulate implies that only certain orbits are permitted, Bohr needed a rule for which orbits were possible. The criterion he chose was that only those orbits are permitted for which the angular momenta have certain discrete values. These values are defined by the relation:

$$mvr = n\frac{h}{2\pi},$$

where h is Planck's constant, and n is a positive integer; that is, n = 1, 2, 3, 4, ... When the possible values of the angular momentum are restricted in this way, the angular momentum is said to be quantized. The integer n which appears in the formula, is called the quantum number. For each value of n there is a stationary state.

See TG, p. 89 for note on angular momentum

With his two postulates and his choice of the permitted stationary states, Bohr was able to calculate additional properties of the stationary states: the radius of each permitted orbit, the speed of the electron in the orbits, and the total energy of the electron in the orbit; this energy is the energy of the stationary state.

The results that Bohr obtained may be summarized in three

In 1908, Rutherford received a Nobel prize. Interestingly enough, the award was in chemistry and not in physics.

simple formulas. The radius of an orbit with quantum number n is given by the expression:

$$r_n = n^2 r_1,$$

The value for r_1 agreed with the size of the hydrogen molecule obtained from the kinetic theory of gases.

where r_1 is the radius of the first orbit (the orbit for n = 1), and has the value 5.29×10^{-9} cm or 5.29×10^{-11} m.

The speed of the electron in the orbit with quantum number n is:

$$v_n = \frac{1}{n} v_1,$$

where v_1 is the speed of the electron in the first orbit, and has the value 2.2×10^8 cm/sec or 2.2×10^6 m/sec. (5×10^6 mph)

The energy of the electron in the orbit with quantum number n is:

$$E_n = \frac{1}{n^2} E_1,$$

where E_1 is the energy of the electron in the first orbit, and has the value -13.6 electron volts or -21.76×10^{-19} joule.

It may seem strange to you that the energy is written with a negative value. Recall that, since it is only changes in energy that can be measured, the zero level for energy can be defined in any way that is convenient. It is customary to define the potential energy of an electron in the field of a nucleus so that it is zero at a very large (or infinite) distance from the nucleus. An energy of zero implies, then, that the electron is just free from the nucleus. Positive values of energy imply that the electron is free of the nucleus and has kinetic energy besides. Negative values of energy imply that the electron is bound to the atom; the more negative, the less the total energy. The lowest energy possible for an electron in a hydrogen atom is -13.6 eV, for which n = 1. This is called the "ground" state.

According to the formula for r_n, the first Bohr orbit has the smallest radius, with n = 1. Higher values of n correspond to orbits that have larger radii. Although the higher orbits are spaced increasingly far apart, the force field of the nucleus falls off rapidly, so the work required to move out to the next orbit actually becomes smaller. Therefore the jumps from one energy level to the next become smaller and smaller at higher energies.

Q11 What was the main evidence that an atom could exist only in certain energy states?

Q12 What reason did Bohr give for the atom existing only in certain energy states?

1901	Wilhelm Röntgen (Ger)—discovery of x-rays.
1902	H. A. Lorentz and P. Zeeman (Neth)—influence of magnetism on radiation.
1903	A. H. Becquerel (Fr)—discovery of spontaneous radioactivity. Pierre and Marie Curie (Fr)—work on rays first discovered by Becquerel.
1904	Lord Rayleigh (Gr Brit)—density of gases and discovery of argon.
1905	Philipp Lenard (Ger)—work on cathode rays.
1906	J. J. Thomson (Gr Brit)—conduction of electricity by gases.
1907	Albert A. Michelson (US)—optical precision instruments and spectroscopic and metrological investigations.
1908	Gabriel Lippmann (Fr)—color photography by interference.
1909	Guglielmo Marconi (Ital)—and Ferdinand Braun (Ger)—development of wireless telegraphy.
1910	Johannes van der Waals (Neth)—equation of state for gases and liquids.
1911	Wilhelm Wien (Ger)—laws governing the radiation of heat.
1912	Nils Gustaf Dalen (Swed)—automatic gas regulators for lighthouses and buoys.
1913	Kamerlingh Onnes (Neth)—low temperature and production of liquid helium.
1914	Max von Laue (Ger)—diffraction of Röntgen rays by crystals.
1915	W. H. and W. L. Bragg (Gr Brit)—analysis of crystal structure by Röntgen rays.
1916	No award.
1917	Charles Glover Barkla (Gr Brit)—discovery of Röntgen radiation of the elements.
1918	Max Planck (Ger)—discovery of energy quanta.
1919	Johannes Stark (Ger)—discovery of the Doppler effect in canal rays and the splitting of spectral lines in electric fields.
1920	Charles-Édouard Guillaume (Switz)—discovery of anomalies in nickel steel alloys.
1921	Albert Einstein (Ger)—for services in theoretical physics and especially for his discovery of the law of the photoelectric effect.
1922	Niels Bohr (Den)—atomic structure and radiation.
1923	Robert Andrews Millikan (US)—elementary charge of electricity and photoelectric effect.
1924	Karl Siegbahn (Swed)—field of x-ray spectroscopy.
1925	James Franck and Gustav Hertz (Ger)—laws governing the impact of an electron upon an atom.
1926	Jean Baptiste Perrin (Fr)—discontinuous structure of matter and especially for his discovery of sedimentation equilibrium.
1927	Arthur Compton (US)—discovery of effect named after him. C. T. R. Wilson (Gr Brit)—method of making paths of electrically charged particles visible by condensation of vapor.
1928	Owen Williams Richardson (Gr Brit)—thermionic phenomena and discovery of effect named after him.
1929	Louis-Victor de Broglie (Fr)—discovery of wave nature of electrons.
1930	Sir Chandrasekhara V. Raman (Ind)—scattering of light and effect named after him.
1931	No award.
1932	Werner Heisenberg (Ger)—quantum mechanics leading to discovery of allotropic forms of hydrogen.
1933	Erwin Schrödinger (Ger) and P. A. M. Dirac (Gr Brit)—new productive forms of atomic theory.
1934	No award.
1935	James Chadwick (Gr Brit)—discovery of the neutron.
1936	Victor Franz Hess (Aus)—cosmic radiation. Carl David Anderson (US)—discovery of the positron.

Nobel Prize winners in Physics.

ALFR. NOBEL
NAT. MDCCCXXXIII OB. MDCCCXCVI

The Nobel Prize

Alfred Bernhard Nobel (1833-1896), a Swedish chemist, was the inventor of dynamite. As a result of his studies of explosives, Nobel found that when nitroglycerine (an extremely unstable chemical) was absorbed in an inert substance it could be used safely as an explosive. This combination is dynamite. He also invented other explosives (blasting gelatin and ballistite) and detonators. Nobel was primarily interested in the peaceful uses of explosives, such as mining, road building and tunnel blasting, and he amassed a large fortune from the manufacture of explosives for these applications. Nobel abhorred war and was conscience-stricken by the military uses to which his explosives were put. At his death, he left a fund of some $315 million to honor important accomplishments in science, literature and international understanding. Prizes were established to be awarded each year to persons who have made notable contributions in the fields of physics, chemistry, medicine or physiology, literature and peace. The first Nobel Prizes were awarded in 1901. Since then, men and women from about 30 countries have received prizes. The Nobel Prize is generally considered the most prestigious prize in science.

Summary 19.6
1. The emission of radiation from atoms accompanies a transition of an electron into a lower energy state.
2. From the postulates of Bohr, the formulas for the spectral series of hydrogen can be derived.

19.6 The Bohr theory: the spectral series of hydrogen. Bohr could now use his model to derive the Balmer formula by applying his second postulate: the radiation emitted or absorbed in a transition has a frequency f determined by the relation

$$hf = E_i - E_f .$$

If n_f is the quantum number of the final state and n_i is the quantum number of the initial state, then according to the E_n formula we have

$$E_f = \frac{E_1}{{n_f}^2} \qquad \text{and} \qquad E_i = \frac{E_1}{{n_i}^2} .$$

The frequency of radiation emitted or absorbed when the atom goes from the initial state to the final state is therefore determined by the equation

$$hf = \frac{E_1}{{n_i}^2} - \frac{E_1}{{n_f}^2} .$$

Balmer's original formula (p. 70) was written in terms of wavelength instead of frequency. The relation between frequency and wavelength was given in Unit 4: the frequency is equal to the speed of the wave divided by its wavelength,

$$f = \frac{c}{\lambda}$$

If we substitute c/λ for f in the equation above, and then divide both sides by the constant hc (Planck's constant times the speed of light), we obtain the equation:

$$\frac{1}{\lambda} = \frac{E_1}{hc}\left[\frac{1}{{n_i}^2} - \frac{1}{{n_f}^2}\right] .$$

According to Bohr's model, then, this equation gives the wavelength λ of the radiation that will be emitted or absorbed when the state of a hydrogen atom changes from n_i to n_f. How does this formula compare with Balmer's formula? The Balmer formula was given on page 70:

$$\frac{1}{\lambda} = R_H\left[\frac{1}{2^2} - \frac{1}{{n_i}^2}\right] .$$

We see at once that the equation derived from the Bohr model is exactly the same as Balmer's formula if:

$$n_f = 2 \qquad \text{and} \qquad R_H = -\frac{E_1}{hc} .$$

F42: A new reality

1937	Clinton J. Davisson (US)—and George P. Thomson (Gr Brit)—experimental diffraction of electrons by crystals.
1938	Enrico Fermi (Ital)—new radioactive elements produced by neutron irradiation and nuclear reactions by slow neutrons.
1939	Ernest O. Lawrence (US)—cyclotron and its use in regard to artificial radioactive elements.
1940	No award
1941	No award
1942	No award
1943	Otto Stern (Ger)—molecular ray method and magnetic moment of the proton.
1944	Isidor Isaac Rabi (US)—resonance method for magnetic properties of atomic nuclei.
1945	Wolfgang Pauli (Aus)—exclusion or Pauli principle.
1946	P. W. Bridgman (US)—high pressure physics.
1947	Sir Edward V. Appleton (Gr Brit)—physics of the upper atmosphere and discovery of so-called Appleton layers.
1948	Patrick M. S. Blackett, (Gr Brit)—development of Wilson cloud chamber and discoveries in nuclear physics and cosmic rays.
1949	Hideki Yukawa (Japan)—prediction of mesons and theory of nuclear forces.
1950	Cecil Frank Powell (Gr Brit)—photographic method of studying nuclear processes and discoveries regarding mesons.
1951	Sir John D. Cockcroft and Ernest T. S. Walton (Gr Brit)—transmutation of atomic nuclei by artificially accelerated atomic particles.
1952	Felix Bloch (Switz) and Edward M. Purcell (US)—nuclear magnetic precision measurements.
1953	Frits Zernike (Neth)—phase-contrast microscope.
1954	Max Born (Ger)—statistical interpretation of wave functions, and Walter Bothe (Ger)—coincidence method for nuclear reactions and cosmic rays.
1955	Willis E. Lamb (US)—fine structure of hydrogen spectrum, and Polykarp Kusch (US)—precision determinations of magnetic moment of electron.
1956	William Shockley, John Bardeen and Walter Houser Brattain (US)—researches on semiconductors and their discovery of the transistor effects.
1957	Chen Ning Yang and Tsung Dao Lee (Chin)—investigation of laws of parity, leading to discoveries regarding the elementary particles.
1958	Pavel A. Cerenkov, Il'ya M. Frank and Igor E. Tamm (USSR)—discovery and interpretation of the Cerenkov effect.
1959	Emilio G. Segrè and Owen Chamberlain (US)—discovery of the antiproton.
1960	Donald A. Glaser (US)—invention of bubble chamber.
1961	Robert Hofstadter (US)—electron scattering in atomic nuclei. Rudolf Ludwig Mössbauer (Ger)—resonance absorption of γ-radiation and discovery of effect which bears his name.
1962	Lev D. Landau (USSR)—theories for condensed matter, especially liquid helium.
1963	Eugene P. Wigner (US)—theory of the atomic nucleus and elementary particles. Maria Goeppert-Mayer (US) and J. Hans D. Jensen (Ger)—nuclear shell structure.
1964	Charles Townes (US), Alexander Prokhorov and Nikolay Basov (USSR)—development of maser.
1965	S. Tomonaga (Japan), Julian Schwinger and Richard Feynman (US)—quantum electrodynamics and elementary particles.
1966	Alfred Kastler (Fr)—new optical methods for studying properties of atom.
1967	Hans Bethe (US)—nuclear physics and theory of energy production in the sun.

Nobel Prize winners in Physics.

Orbits are not drawn precisely to scale — r_n increases as n^2

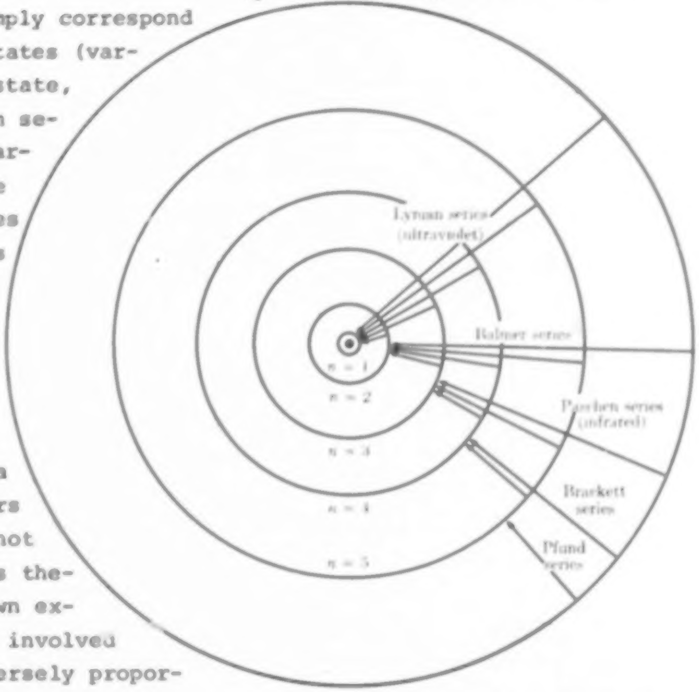


Fig. 19.10 Possible transitions of an electron in the Bohr model of the hydrogen atom.

SG 19.11

T39: Energy levels — Bohr theory

All the lines in the Balmer series simply correspond to transitions from various initial states (various values of n_i) to the same final state, $n_f = 2$. Similarly, lines of the Lyman series correspond to transitions from various initial states to the final state $n_f = 1$; the lines of the Paschen series correspond to transitions from various initial states to the final state $n_f = 3$, etc. (see Table 19.2). The general scheme of possible transitions is shown in Fig. 19.10.

The Bohr formula, for hydrogen, agrees exactly with the Balmer formula as far as the dependence on the numbers n_f and n_i is concerned. But this is not surprising, since Bohr constructed his theory in such a way as to match the known experimental results. Any theory which involved stationary states whose energy is inversely proportional to the square of a quantum number n would do as well as this. Of course any such theory would have to rely in some way on the idea that radiation is quantized and that the electron has stationary states in the atom. Not only did Bohr's model lead to correct dependence on n_f and n_i, but more remarkably, the value of the constant came out right. The Rydberg constant R_H, which had previously been just an experimentally determined constant, was now shown to depend on the mass and charge of the electron, on Planck's constant and on the speed of light.

When the Bohr theory was proposed, in 1913, only the Balmer and Paschen series for hydrogen were known. The theory suggested that additional series should exist. The experimental search for these series yielded the Lyman series in the ultraviolet portion of the spectrum (1916), the Brackett series (1922), and the Pfund series (1924). In each series the measured frequencies of the lines were found to be those predicted by the theory. Thus, the theory not only correlated known information about the spectrum of hydrogen, but also predicted hitherto unknown series of lines in the spectrum.

The scheme shown in Fig. 19.10 is useful, but it also has the danger of being too specific. For instance, it leads us to visualize the emission of radiation in terms of

The possible excitations of Mercury from the ground state are shown at the far right. However, unlike the hydrogen atom, transitions between these levels are not possible – they are "forbidden" by the "selection rules" which arise from other quantum properties of the atom. (Similar restric-

19.6

"jumps" of electrons between orbits. But we cannot actually detect an electron moving in an orbit, nor can we see an electron "jump" from one orbit to another. A second way of presenting the results of Bohr's theory was suggested, which yields the same facts but does not commit us too closely to a picture of orbits. This new scheme is shown in Fig. 19.11. It focusses attention on the possible energy states, which are all given by the formula, $E_n = \frac{1}{n^2} E_1$. In terms of this mathematical model, the atom is normally unexcited, its energy then being E_1, or $-22 \cdot 10^{-19}$ joules. Absorption of energy can place the atoms in an excited state, with a larger energy. The excited atom is then ready to emit light with a consequent reduction in energy. But the energy absorbed or emitted must always shift the energy of the atom to one of the values specified by the E_n formula. We may thus, if we wish, represent the hydrogen atom by means of the energy-level diagram shown on the left.

There is a continuum of possible energy values once the electron is free of the atom.

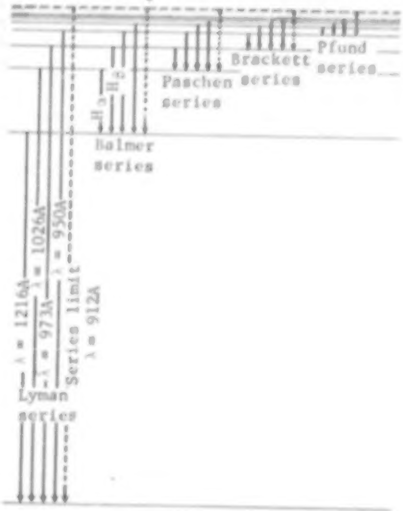


Fig. 19.11 Energy-level diagram for the hydrogen atom. The energy units are 10^{-19} joules.

D57: Absorption

Q13 **Balmer had predicted accurately the other spectral series in hydrogen thirty years before Bohr did. Why is Bohr's prediction considered more important?**

19.7 Stationary states of atoms: the Franck-Hertz experiment.

The success of the Bohr theory in accounting for the spectrum of hydrogen raised the question: can experiments show directly that atoms have only certain discrete energy states? In other words, are there really gaps between the energies that an atom can have? A famous experiment in 1914, by the German Physicists James Frank and Gustav Hertz, showed the existence of these discrete energy states.

James Franck (1882-1946) and Gustav Hertz (1887-) won a Nobel Prize for their work in 1925. In the 1930's they both were dismissed from their university posts because they were of Jewish descent. Franck fled to the United States and worked on the atomic bomb during World War II. He tried to have the bomb's power demonstrated before an international group in a test instead of in the destruction of Japanese cities. Hertz chose to remain in Germany. He survived in one of the concentration camps that were liberated by Russian forces in 1945.

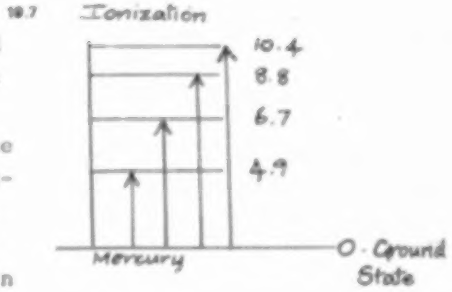
Franck and Hertz bombarded atoms with electrons (from an electron gun) and were able to measure the energy lost by electrons in collisions with atoms. They could also determine the energy gained by atoms in these collisions. Their work was very ingenious, but it is too complex to describe and interpret in detail in this course. We shall therefore give here a somewhat oversimplified account of their experiments.

In their first experiment, Franck and Hertz bombarded mercury atoms in mercury vapor contained in a chamber at very low pressure. Their experimental procedure was equivalent to measuring the kinetic energy of electrons leaving the

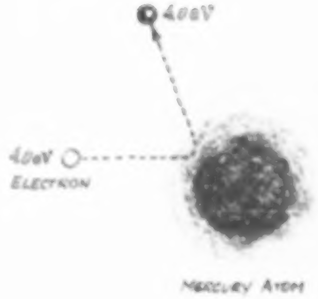
Summary 19.7

1. The existence of stationary energy states, a vital part of the Bohr theory, was experimentally verified by Franck and Hertz. In a study of the energy exchanges in collisions between electrons and various atoms, they found that only discrete quantities of energy could be transferred.

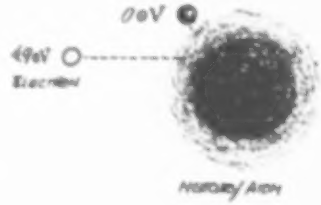
tions apply to the hydrogen atom, but because the energy sublevels are usually so close, the diagram is acceptably correct.)



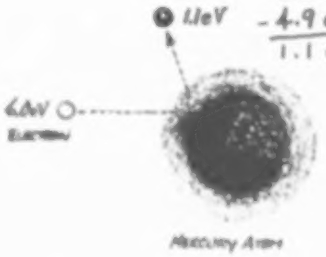
electron gun and the kinetic energy of electrons that had passed through the mercury vapor. The only way electrons could lose energy was in collisions with mercury atoms. Franck and Hertz found that when the kinetic energy of the electrons leaving the electron gun was very small, for example, about 1 eV, the electrons that passed through the mercury vapor had almost exactly the same energy as they had on leaving the gun. This result could be explained in the following way. A mercury atom is several hundred thousand times more massive than an electron. At low electron energies the electron just bounces off a mercury atom, much as a golf ball thrown at a bowling ball would bounce off it. A collision of this kind is called an "elastic" collision. In an elastic collision, the mercury atom (bowling ball) takes up only an extremely small part of the kinetic energy of the electron (golf ball). The electron loses practically none of its kinetic energy.



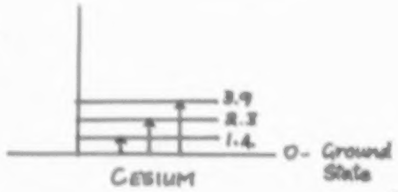
When the energy of the bombarding electrons was raised to 5 eV, there was a dramatic change in the experimental results. An electron that collided with a mercury atom lost almost exactly 4.9 electron-volts of energy. When the electron energy was increased to about 6 electron-volts, an electron still lost 4.9 electron-volts of energy in a collision with a mercury atom. The electron had just 1.1 eV of energy after passing through the mercury vapor. These results indicated that a mercury atom cannot accept less than 4.9 eV of energy; and that when it is offered somewhat more, for example, 5 or 6 eV, it still can accept only 4.9 eV. This energy cannot go into kinetic energy of the mercury atom because of the relatively enormous mass of the atom as compared with that of an electron. Hence, Franck and Hertz concluded that the 4.9 eV of energy is added to the internal energy of the mercury atom—that the mercury atom has a permitted or stationary state with energy 4.9 eV greater than that of the lowest energy state. They also concluded that there is no state with an energy in between.



What happens to this 4.9 eV of additional internal energy? According to the Bohr theory, if the mercury atom has a state with energy 4.9 eV greater than that of the lowest state, this amount of energy should be emitted in the form of electromagnetic radiation when the atom returns to its lowest state. Franck and Hertz looked for this radiation with a spectroscope, and found it. They observed a spectrum line at a wavelength of 2535 Å, a line that was known in the emission spectrum of mercury. The wavelength corresponds to a frequency f that is equivalent to an energy, hf, of 4.9 eV.



$$\begin{array}{r} 6.0 \text{ ev} \\ -\ 4.9 \text{ ev} \\ \hline 1.1 \text{ ev} \end{array}$$



Use the diagram for Cesium to present other examples. →
(see conditions for Hg. above)

F43: Franck-Hertz experiment

19.7 This result showed that the mercury atoms had indeed gained 4.9 eV of energy in their collisions with the electrons.

SG 19.12

Any student interested in duplicating the Franck-Hertz experiment (for a science fair project, for example) should see Novel Experiments in Physics, American Institute of Physics, 1964 p. 305. This is a difficult experiment.

Later experiments showed that mercury atoms could also gain other, sharply defined amounts of energy when bombarded with electrons, for example, 6.7 eV and 10.4 eV. In each case radiation was emitted that corresponded to lines in the spectrum of mercury. Experiments have also been made on many other elements besides mercury; in each case analogous results were obtained. The electrons always lost energy, and the atoms always gained energy in sharply defined amounts. Each type of atom studied was found to have discrete energy states. The amounts of energy gained by the atoms in collisions with electrons could always be correlated with spectrum lines. The existence of discrete "permitted" or "stationary" states of atoms predicted by the Bohr theory of atomic spectra was thus verified by direct experiment. This verification was considered to provide strong confirmation of the validity of the Bohr theory.

Q14 How much kinetic energy will an electron have after a collision with a mercury atom if its kinetic energy before collision is (a) 4.0 eV? (b) 5.0 eV? (c) 7.0 eV?

Summary 19.8
The Bohr theory and the periodic properties of the elements together combine to suggest an image of electrons in the atom grouped into shells and subshells of different energies. Thus, a physical basis for understanding the periodic properties (chemical and physical) of elements resulted from Bohr's work.

19.8 The periodic table of the elements. In the Rutherford-Bohr model, the atoms of the different elements differ in the charge and mass of the nucleus, and in the number and arrangement of the electrons about the nucleus. As for the arrangement of the electrons, Bohr came to picture the electronic orbits as on the next page, though not as a series of concentric rings in one plane but as tracing out patterns in three dimensions. For example, the orbits of the two electrons of He in the normal state are indicated as circles in planes inclined at about 60° with respect to each other. In addition to circular orbits, elliptical ones with the nucleus at one focus are also possible.

The German physicist Arnold Sommerfeld (1868-1951) in 1915 improved the Bohr atomic model by introducing elliptical orbits so as to allow more freedom in choosing the "permitted" orbits. The eccentricity of the ellipse was also quantized and associated with another quantum number.

Bohr found a way of correlating his model with the periodic table of the elements and the periodic law. He suggested that the chemical and physical properties of an element depend on how the electrons are arranged around the nucleus. He also indicated how this might come about. He regarded the electrons in an atom as grouped into shells. Each shell can contain not more than a certain number of electrons. The chemical properties are related to how nearly full or empty a shell is. For example, full shells are associated with chemical stability, and in the inert gases the electron shells are completely filled.

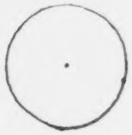
SG 19.13

To relate the Bohr model of atoms with their chemical properties we may begin with the observation that the elements hydrogen (Z = 1) and lithium (Z = 3) are somewhat alike chemically. Both have valences of 1. Both enter into compounds of analogous types, for example hydrogen chloride, HCl, and lithium chloride, LiCl. Furthermore, there are some similarities in their spectra. All this suggests that the lithium atom resembles the hydrogen atom in some important respects. Bohr conjectured that two of the three electrons of the lithium atom are relatively close to the nucleus, in orbits like those pertinent to the helium atom, while the third is in a circular or elliptical orbit outside the inner system. Since this inner system consists of a nucleus of charge (+) $3q_e$ and two electrons each of charge (-) q_e, its net charge is (+) q_e. Thus the lithium atom may be roughly pictured as having a central core of charge (+) q_e, around which one electron revolves, somewhat as for a hydrogen atom.

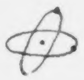
Helium (Z = 2) is a chemically inert element, belonging to the family of noble gases. So far no one has been able to form compounds from it. These properties indicated that the helium atom is highly stable, having both of its electrons closely bound to the nucleus. It seemed sensible to regard both electrons as moving in the same innermost shell around the nucleus when the atom is unexcited. Moreover, because of the stability and the chemical inertness of the helium atom, we may reasonably assume that this shell cannot accommodate more than two electrons. This shell is called the K-shell. The single electron of hydrogen is also said to be in the K-shell when the atom is unexcited. For lithium, two electrons are in the K-shell, filling it to capacity, and the third electron starts a new one, called the L-shell. To this single outlying and loosely bound electron must be ascribed the strong chemical affinity of lithium for oxygen, chlorine and many other elements.

Sodium (Z = 11) is the next element in the periodic table that has chemical properties similar to those of hydrogen and lithium, and this suggests that the sodium atom also is hydrogen-like in having a central core about which on electron revolves. Moreover, just as lithium follows helium in the periodic table, so does sodium follow another noble gas, neon (Z = 10). For the neon atom, we may assume that 2 of its 10 electrons are in the first (K) shell, and that the remaining 8 electrons are in the second (L) shell. Because of the great chemical inertness and stability of neon, these 8 electrons may be expected to fill the L-shell to capacity.

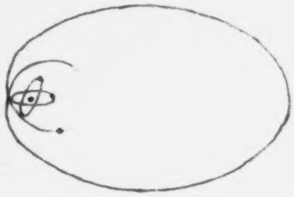
The sketches below are based on diagrams Bohr used in his university lectures.



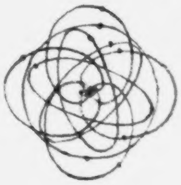
HYDROGEN (Z=1)



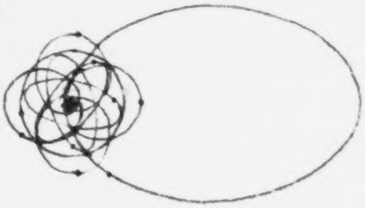
HELIUM (Z=2)



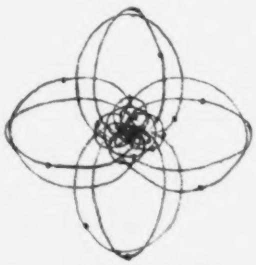
LITHIUM (Z=3)



NEON (Z=10)



SODIUM (Z=11)



ARGON (Z=18)

These two pages will be easier to study if you refer to the table of the elements and the periodic table in Chapter 18.

For sodium, then, the eleventh electron must be in a third shell, which is called the M-shell. Passing on to potassium (Z = 19), the next alkali metal in the periodic table, we again have the picture of an inner core and a single electron outside it. The core consists of a nucleus with charge (+) $19q_e$ and 2, 8, and 8 electrons occupying the K-, L-, and M-shells, respectively. The 19th electron revolves around the core in a fourth shell, called the N-shell. The atom of the noble gas, argon with Z = 18 just before potassium in the periodic table, again represents a distribution of electrons in a tight and stable pattern, with 2 in the K-, 8 in the L-, and 8 in the M-shell.

These qualitative considerations have led us to a consistent picture of electrons in groups, or shells, around the nucleus. The arrangement of electrons in the noble gases can be taken to be particularly stable, and each time we encounter a new alkali metal in Group I of the periodic table, a new shell is started with a single electron around a core which resembles the pattern for the preceding noble gas. We may expect that this outlying electron will easily come loose under the attraction of neighboring atoms, and this corresponds with the facts. The elements lithium, sodium and potassium belong to the group of alkali metals. In compounds or in solution (as in electrolysis) they may be considered to be in the form of ions such as Li^+, Na^+ and K^+, each with one positive net charge $(+)q_e$. In the atoms of these elements, the outer electron is relatively free to move about. This property has been used as the basis of a theory of electrical conductivity. According to this theory, a good conductor has many "free" electrons which can form a current under appropriate conditions. A poor conductor has relatively few "free" electrons. The alkali metals are all good conductors. Elements whose electron shells are filled are very poor conductors because they have no "free" electrons.

About half of the elements can be classified into eight well-defined families. The remaining elements, which are not included in these eight families, are all metallic in character.

Turning now to Group II of the periodic table, we would expect those elements that follow immediately after the alkali metals to have atoms with two outlying electrons. For example, beryllium (Z = 4) should have 2 electrons in the K-shell, thus filling it, and 2 in the L-shell. If the atoms of all these elements have two outlying electrons, they should be chemically similar, as indeed they are. Thus, calcium and magnesium, which belong to this group, should easily form ions such as Ca^{++} and Mg^{++}, each with two positive charges, $(+)2q_e$, and this is also found to be true.

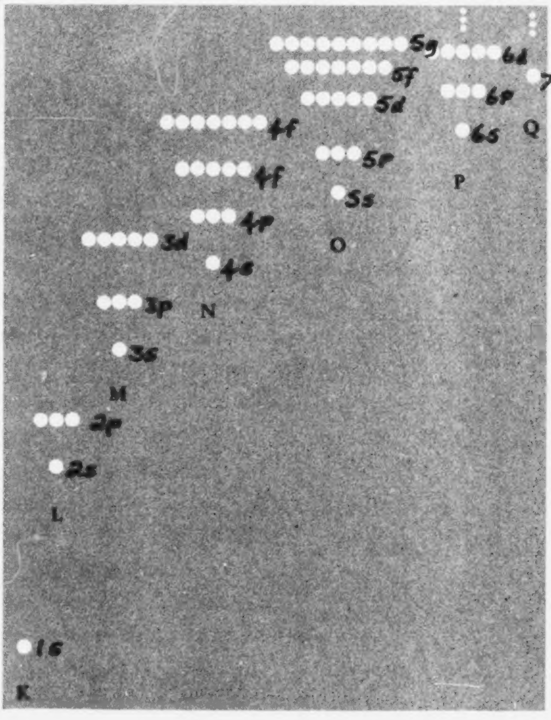
In 1925 Wolfgang Pauli (Austria, Switzerland, 1900-1958) announced the principle 'any given quantum orbit in an atom can be occupied by no more than two electrons.' 19.8

As a final example, consider those elements that immediately precede the noble gases in the periodic table. For example, fluorine atoms (Z = 9) should have 2 electrons filling the K-shell but only 7 electrons in the L-shell, which is one less than enough to fill it. If a fluorine atom should capture an additional electron, it should become an ion F^- with one negative charge. The L-shell would then be filled, as it is for neutral neon (Z = 10), and thus we would expect the F^- ion to be stable. This prediction is in accord with observation. Indeed, all the elements immediately preceding the inert gases in the periodic table tend to form stable singly charged negative ions in solution. In the solid state, we would expect these elements to be lacking in free electrons, and all of them are in fact poor conductors of electricity.

This was called the "exclusion principle" by Pauli himself. He was led to his principle through the study of atomic spectra. It was Uhlenbeck and Goudsmit, about nine months later, who introduced a fourth quantum number called the electron spin quantum number 's'. Stern and Gerlach demonstrated the existence of electron spin in their famous experiment in 1921.

Altogether there are seven main shells, K, L, M, ... Q, and further analysis shows that all but the first are divided into subshells. Thus the first shell K is one shell without substructure, the second shell L consists of two subshells, and so on. The first subshell in any shell can always hold up to 2 electrons, the second up to 6, the third up to 10, the fourth up to 14, and so on. Electrons that are in different subsections of the same shell in general differ very little in energy as compared with electrons that are in different shells. For all the elements up to and including argon (Z = 18), the buildup of electrons proceeds quite simply. Thus the argon atom has 2 electrons in the K-shell, 8 in the L-shell, then 2 in the first M-subshell and 6 in the second M-subshell. But after argon, there may be electrons in an outer shell before an inner one is filled. This complicates the scheme somewhat but still allows it to be consistent. The arrangement of the electrons in any unexcited atom is always the one that provides greatest stability for the whole atom. According to this model, chemical phenomena generally involve only the outermost electrons of the atoms.

These different subshells are associated with the quantum number associated with Sommerfeld's elliptical orbits. In Chapter 20, they can be related to the different probability distributions.
The subshells in a main shell are designated (in increasing energy) by the letters s, p, d, f, ...



Relative energy levels of electron states in atoms. Each circle represents a state which can be occupied by 2 electrons.

SG 19.14
SG 19.15

The first deviation from a simple progression occurs for the 19th electron; it goes into the 4s orbital instead of the expected 3d. The energy of the 4s orbital is lower than that of the 3d orbital.

T35: Periodic table

19.8

Bohr carried through a complete analysis along these lines and, in 1921, proposed the form of the periodic table shown in Fig. 19.12. This table was the result of physical theory and offered a fundamental physical basis for understanding chemistry. This was another triumph of the Bohr theory.

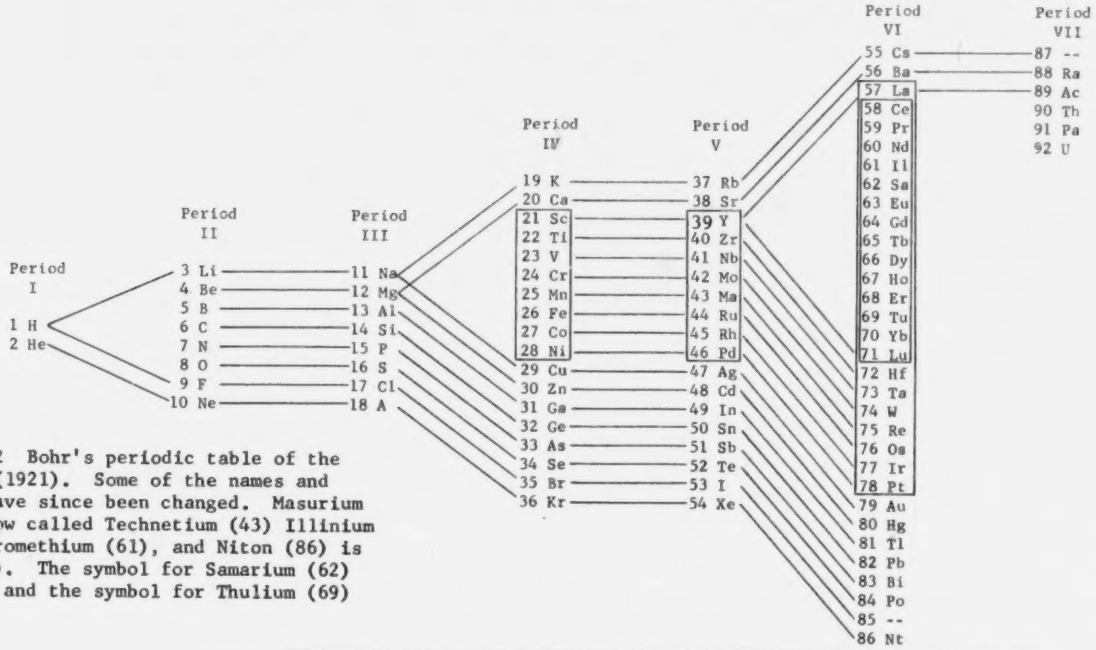

Period I
1 H
2 He
Period II
3 Li
4 Be
5 B
6 C
7 N
8 O
9 F
10 Ne
Period III
11 Na
12 Mg
13 Al
14 Si
15 P
16 S
17 Cl
18 A
Period IV
19 K
20 Ca
21 Sc
22 Ti
23 V
24 Cr
25 Mn
26 Fe
27 Co
28 Ni
29 Cu
30 Zn
31 Ga
32 Ge
33 As
34 Se
35 Br
36 Kr
Period V
37 Rb
38 Sr
39 Y
40 Zr
41 Nb
42 Mo
43 Ma
44 Ru
45 Rh
46 Pd
47 Ag
48 Cd
49 In
50 Sn
51 Sb
52 Te
53 I
54 Xe
Period VI
55 Cs
56 Ba
57 La
58 Ce
59 Pr
60 Nd
61 Il
62 Sa
63 Eu
64 Gd
65 Tb
66 Dy
67 Ho
68 Er
69 Tu
70 Yb
71 Lu
72 Hf
73 Ta
74 W
75 Re
76 Os
77 Ir
78 Pt
79 Au
80 Hg
81 Tl
82 Pb
83 Bi
84 Po
85 --
86 Nt
Period VII
87 --
88 Ra
89 Ac
90 Th
91 Pa
92 U

Fig. 19.12 Bohr's periodic table of the elements (1921). Some of the names and symbols have since been changed. Masurium (43) is now called Technetium (43) Illinium (61) is Promethium (61), and Niton (86) is Radon (86). The symbol for Samarium (62) is now Sm and the symbol for Thulium (69) is Tm.

Q15 Why do the next heavier elements after the noble gases easily become positively charged?

Summary 19.9
Although the Bohr theory accounted for many of the physical and chemical properties of the elements, it had its limitations and shortcomings. The Bohr theory was a hybrid-a mixture of classical and quantum ideas. A new theory, based fundamentally on quantum concepts, was needed.

19.9 The failure of the Bohr theory and the state of atomic theory in the early 1920's. In spite of the successes achieved with the Bohr theory in the years between 1913 and 1924, serious problems arose for which the theory proved inadequate. Although the Bohr theory accounted for the spectra of atoms with a single electron in the outermost shell, serious discrepancies between theory and experiment appeared in the spectra of atoms with two electrons in the outermost shell. Indeed the theory could not account in any satisfactory way for the spectra of elements whose atoms have more than one electron in the outermost shell. It was also found experimentally that when a sample of an element is in an electric or magnetic field, its emission spectrum shows additional lines. For example, in a magnetic field each line is split into several lines. The Bohr theory could not account in a quantitative way for the observed splitting. Further, the theory supplied no method for predicting the relative intensities of spectral lines. These intensities depend on the numbers of atoms in a sample that undergo transitions among

During World War II, Niels Bohr had to flee his laboratory in Denmark. Two of his colleagues gave him their gold Nobel Prize medals, which he dissolved in a jar of acid and left in his office. Believing in the conservation of matter, he was certain that the acid would only dissolve the gold and he knew that the Nazis would never suspect the jar of acid to be valuable. When he returned after the war, he regained the gold from the acid and had the medals recast.

•His doctoral thesis contained a foreshadowing of his theory of the atom, where he points out that classical mechanics is insufficient to explain the structure of matter.

Bohr taking his wife for a ride (1931). The motorcycle was owned by George Gamow.

Niels Bohr (1885-1962) was born in Copenhagen, Denmark and was educated there, receiving his doctor's degree• in physics in 1911. In 1912 he was at work in Rutherford's laboratory in Manchester, England, which was a center of research on radioactivity and atomic structure. Here he developed his theory of atomic structure and atomic spectra. Bohr played an important part in the development of quantum mechanics, in the advancement of nuclear physics, and in the study of the philosophical aspects of modern physics. In his later years he devoted much time to promoting the peaceful uses of atomic and nuclear physics.

In addition to the Nobel Prize, Bohr was given the first Atoms for Peace Award, a $75,000 prize, in 1957.

the stationary states. Physicists wanted to be able to calculate the probability of a transition from one stationary state to another. They could not make such calculations with the Bohr theory.

By the early 1920's it had become clear that the Bohr theory, despite its great successes, had deficiencies and outright failures. It was understood that the theory would have to be revised, or replaced by a new one. The successes of the Bohr theory showed that a better theory of atomic structure would have to account for the existence of stationary states—discrete atomic levels—and would, therefore, have to be based on quantum concepts. Besides the inability to predict certain properties at all, the Bohr theory had two additional shortcomings: it predicted some results that disagreed with experiment; and it predicted others that could not be tested in any known way. Of the former kind were predictions about the spectra of elements with two or three electrons in the outermost electron shells. Of the latter kind were predictions of the details of electron orbits. Details of this latter type could not be observed directly, nor could they be related to any observable properties of atoms such as the lines in the emission spectrum. Planetary theory has very different implications when applied to a planet revolving around the sun, and when applied to an electron in an atom. The precise position of a planet is important, especially if we want to do experiments such as photographing the surface of the moon or of Mars from a satellite. But the calculation of the position of an electron in an orbit is neither useful nor interesting because it has no relation to any experiment physicists have been able to devise. It thus became evident that, in using the Bohr theory, physicists were asking some questions which could not be answered experimentally.

See "The Sea-Captain's Box" in Project Physics Reader 5.

In the early 1920's, physicists began to think seriously about what could be wrong with the basic ideas of the theory. One fact that stood out was that the theory started with a mixture of classical and quantum ideas. An atom was assumed to act in accordance with the laws of classical physics up to the point where these laws didn't work; then the quantum ideas were introduced. The picture of the atom that emerged from this mixture was an inconsistent combination of ideas from classical physics and concepts for which there was no place in classical physics. The orbits of the electrons were determined by the classical, Newtonian laws of motion. But of the many possible orbits, only a small fraction were regarded as possible, and these were assigned by rules that

In March 1913, Bohr wrote to Rutherford enclosing a draft of his first paper on the quantum theory of atomic constitution. On March 20, 1913, Rutherford replied in a letter, the first part of which we quote,

> "Dear Dr. Bohr:
>
> I have received your paper safely and read it with great interest, but I want to look it over again carefully when I have more leisure. Your ideas as to the mode of origin of spectra in hydrogen are very ingenious and seem to work out well; but the mixture of Planck's ideas with the old mechanics makes it very difficult to form a physical idea of what is the basis of it. There appears to me one grave difficulty in your hypothesis, which I have no doubt you fully realize, namely, how does an electron decide what frequency it is going to vibrate at when it passes from one stationary state to the other. It seems to me that you would have to assume that the electron knows beforehand where it is going to stop...."

contradicted classical mechanics. It became evident that a better theory of atomic structure would have to have a more consistent foundation and that the quantum concepts would have to be fundamental, rather than secondary.

The contribution of the Bohr theory may be summarized as follows. It provided partial answers to the questions raised about atomic structure in Chapters 17 and 18. Although the theory turned out to be inadequate it supplied clues to the way in which quantum concepts should be used. It indicated the path that a new theory would have to take. A new theory would have to supply the right answers that the Bohr theory gave and would also have to supply the right answers for the problems the Bohr theory couldn't solve. A successful theory of atomic structure has been developed and has been generally accepted by physicists. It is called "quantum mechanics" because it is built directly on the foundation of quantum concepts; it will be discussed in the next chapter.

SG 19.16
SG 19.17
SG 19.18

Q16 The Bohr model of atoms is widely given in science books. What is wrong with it?

19.1 (a) Suggest experiments to show which of the Fraunhofer lines in the spectrum of sunlight are due to absorption in the sun's atmosphere rather than to absorption by gases in the earth's atmosphere. discussion

(b) How might one decide from spectroscopic observations whether the moon and the planets shine by their own light or by reflected light from the sun? discussion

19.2 Theoretically, how many series of lines are there in the emission spectrum of hydrogen? infinite In all these series, how many lines are in the visible region? four

19.3 The Rydberg constant for hydrogen, R_H, has the value 1.097×10^7/m. Calculate the wavelengths of the lines in the Balmer series corresponding to n = 8, n = 10, n = 12. Compare the values you get with the wavelengths listed in Table 19.1. Do you see any trend in the values?

$n = 8$, $\lambda = 3880$ Å
$n = 10$, $\lambda = 3790$ Å
$n = 12$, $\lambda = 3740$ Å

19.4 (a) As indicated in Fig. 19.5 the lines in one of hydrogen's spectral series are bunched very closely at one end. Does the formula $\frac{1}{\lambda} = R_H\left[\frac{1}{n_f^2} - \frac{1}{n_i^2}\right]$ suggest that such bunching will occur? yes

(b) The series limit must correspond to the last possible line(s) of the series. What value should be taken for n_i in the above equation to compute the wavelength of the series limit? $n_i = \infty$

(c) Compute the series limit for the Lyman, Balmer and Paschen series of hydrogen. 910 Å, 3640 Å, 8180 Å, respectively

(d) Consider a photon with a wavelength corresponding to the series limit of the Lyman series. What energy would it carry? Express the answer in joules and in electron-volts (1 eV = 1.6×10^{-19} J). 21.8×10^{-19} j or 13.6 eV

19.5 In what ways do the Thomson and Rutherford atomic models agree? In what ways do they disagree? discussion

19.6 In 1903, the German physicist; Philipp Lenard (1864-1947), proposed an atomic model different from those of Thomson and Rutherford. He had observed that, since cathode-ray particles can penetrate matter, most of the atomic volume must offer no obstacle to their penetration. In Lenard's model there were no electrons and no positive charges separate from the electrons. His atom was made up of particles called dynamides, each of which was an electric doublet possessing mass. (An electric doublet is a combination of a positive charge and a negative charge very close together.) All the dynamides were supposed to be identical, and an atom contained as many of them as were needed to make up its mass. They were distributed throughout the volume of the atom, but their radius was so small compared with that of the atom that most of the atom was actually empty. discussion

(a) In what ways does Lenard's model agree with those of Thomson and Rutherford? In what ways does it disagree with those models?
(b) Why would you not expect α particles to be scattered through large angles if Lenard's model were valid?

19.7 In a recently published book the author expresses the view that physicists have interpreted the results of the experiments on the scattering of α particles incorrectly. He thinks that the experiments show only that atoms are very small, not that they have a heavy, positively charged nucleus. Do you agree with his view? Why? discussion

19.8 Suppose that the atom and the nucleus are each spherical, that the diameter of the atom is of the order of 1 Å (Ångstrom unit) and that the diameter of the nucleus is of the order of 10^{-12} cm. What is the ratio of the diameter of the nucleus to that of the atom?

10^{-4}

19.9 The nucleus of the hydrogen atom is thought to have a radius of about 1.5×10^{-13} cm. If the nucleus were magnified to 0.1 mm (the radius of a grain of dust), how far away from it would the electron be in the Bohr orbit closest to it?

3.5m

19.10 In 1903 a philosopher wrote,

> The propounders of the atomic view of electricty [disagree with theories which] would restrict the method of science to the use of only such quantities and data as can be actually seen and directly measured, and which condemn the introduction of such useful conceptions as the atom and the electron, which cannot be directly seen and can only be measured by indirect processes.

On the basis of the information now available to you, with which view do you agree; the view of those who think in terms of atoms and electrons, or the view that we must use only such things as can be actually seen and measured?

discussion

19.11 How would you account for the production of the lines in the absorption spectrum of hydrogen by using the Bohr theory?

discussion

19.12 Many substances emit visible radiation when illuminated with ultraviolet light; this phenomenon is an example of fluorescence. Stokes, a British physicist of the nineteenth century, found that in fluorescence the wavelength of the emitted light usually was the same or longer than the illuminating light. How would you account for this phenomenon on the basis of the Bohr theory?

discussion

19.13 In Query 31 of his Opticks, Newton wrote:

> All these things being consider'd, it seems probable to me that God in the beginning formed matter in solid, massy, hard, impenetrable, moveable particles, of such sizes and figures, and with such other properties, and in such proportion to space, as most conduced to the end for which he formed them; and that these primitive particles being solids, are incomparably harder than any porous bodies compounded of them; even so very hard, as never to wear or break in pieces; no ordinary power being able to divide what God himself made one in the first creation. While the particles continue entire, they may compose bodies of one and the same nature and texture in all ages: But should they wear away, or break in pieces, the nature of things depending on them would be changed. Water and earth, composed of old worn particles and fragments of particles, would not be of the same nature and texture now, with water and earth composed of entire particles in the beginning. And therefore, that nature may be lasting, the changes of corporeal things are to be placed only in the various separations and new associations and motions of these permanent particles; compound bodies being apt to break, not in the midst of solid particles, but where those particles are laid together, and only touch in a few points.

Compare what Newton says here about atoms with discussion

a) the views attributed to Leucippus and Democritus concerning atoms (see the prologue to this unit);
b) Dalton's assumptions about atoms (see the end of the prologue to this unit);
c) the Rutherford-Bohr model of the atom.

19.14 Use the chart on p. 91 to explain why atoms of potassium (Z = 19) have electrons in the N shell even though the M shell isn't filled. discussion

19.15 Use the chart on p. 91 to predict the atomic number of the next inert gas after argon. That is, imagine filling the electron levels with pairs of electrons until you reach an apparently stable, or complete, pattern. discussion

Do the same for the next inert gas.

19.16 Make up a glossary, with definitions, of terms which appeared for the first time in this chapter. discussion

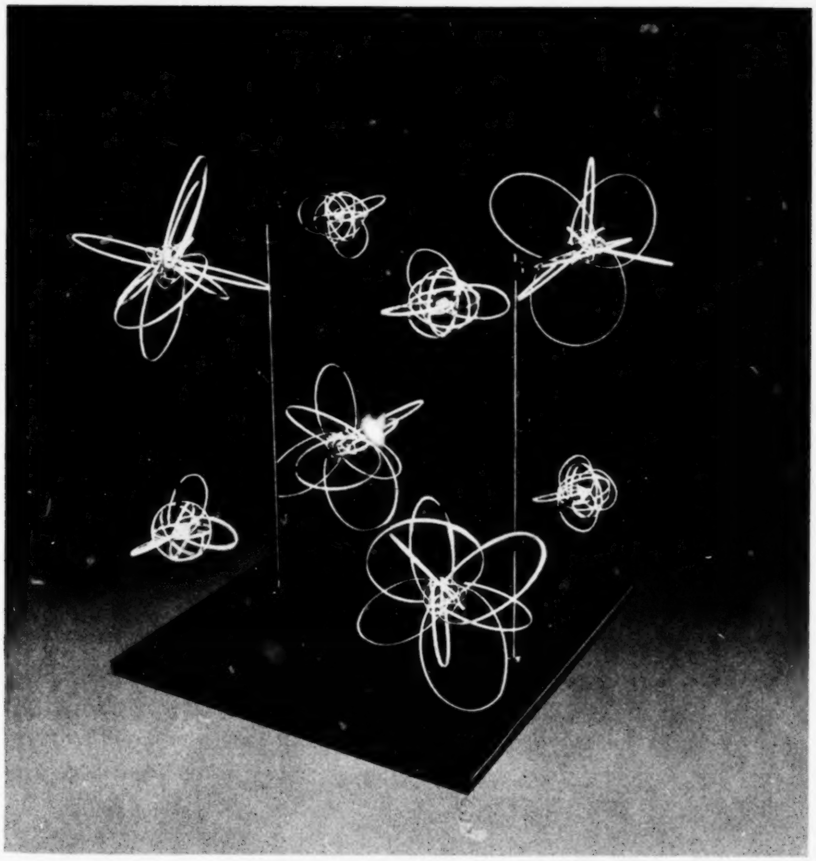
19.17 The philosopher John Locke (1632-1704) proposed a science of human nature which was strongly influenced by Newton's physics. In Locke's atomistic view, elementary ideas are produced by elementary sensory experiences and then drift, collide and interact in the mind. Thus the association of ideas was but a specialized case of the universal interactions of particles. discussion

Does such an "atomistic" approach to the problem of human nature seem reasonable to you? What argument for and against this sort of theory can you think of?

19.18 In a recently published textbook of physics, the following statement is made:

> Arbitrary though Bohr's new postulate may seem, it was just one more step in the process by which the apparently continuous macroscopic world was being analyzed in terms of a discontinuous, quantized, microscopic world. Although the Greeks had speculated about quantized matter (atoms), it remained for the chemists and physicists of the nineteenth century to give them reality. In 1900 Planck found it necessary to quantize the energy of the atomic-sized oscillators responsible for blackbody radiation. In 1905 Einstein quantized the energy of electromagnetic waves. Also, in the early 1900's a series of experiments culminating in Millikan's oil-drop experiment conclusively showed that electric charge was quantized. To this list of quantized entities, Bohr added angular momentum.

a) What other properties or things in physics can you think of that are "quantized?" discussion
b) What properties or things can you think of outside physics that might be said to be "quantized?"



This sculpture is meant to represent the arrangement of sodium and chlorine ions in a crystal of common salt. Notice that the outermost electrons of the sodium atoms have been lost to the chlorine atoms, leaving sodium ions with completed K and L shells and chlorine ions with completed K, L and M shells.

Chapter 20 Some Ideas From Modern Physical Theories

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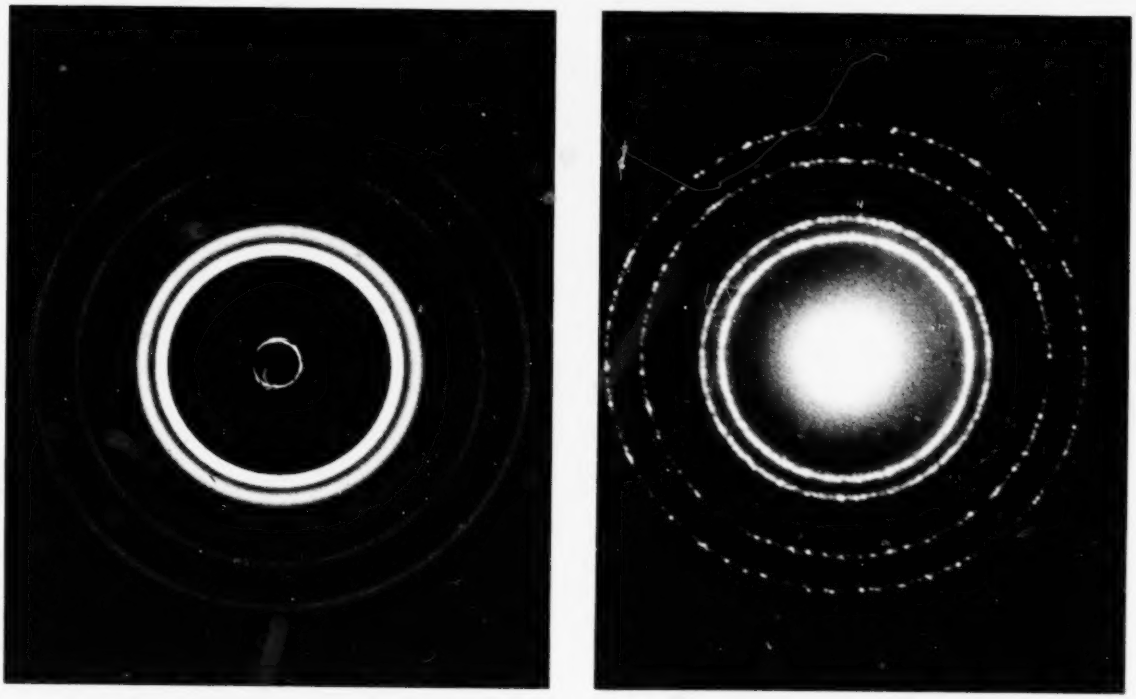
"Thus grew the tale of wonderland:
Thus slowly, one by one,
Its quaint events were hammered out—"
Lewis Carroll,
Introduction to 'Alice in Wonderland'

The diffraction pattern on the left was made by a beam of x rays passing through thin aluminium foil. The diffraction pattern on the right was made by a beam of electrons passing through the same foil.



The em wavelength of the x rays was the same as the de Broglie wavelength of the electrons.

20.1 Some results of relativity theory. Progress in atomic and nuclear physics has been based on two great revolutions in physical thought: quantum theory and relativity. In Chapters 18 and 19 we saw how quantum theory entered into atomic physics. The further development of quantum theory, quantum mechanics, will be the main subject of this chapter. But we cannot get into quantum mechanics without learning something about relativity. Some of the results of the relativity theory are needed to understand certain phenomena of atomic physics which are basic to quantum mechanics. These results will also be essential to our treatment of nuclear physics in Unit 6. We shall, therefore, devote this section to a brief discussion of the theory of relativity, introduced by Einstein in 1905—the same year in which he published the theory of the photoelectric effect.

The theory of relativity ties together ideas and experimental information that have been touched on earlier in this course. One important piece of information involves the speed of light. Measurements showed a remarkable and surprising result: the speed of light in vacuum (free space) is independent of any motion of the source of the light or of the person making the measurement. The result is always the same, 3.0×10^8 m/sec, regardless of whether the measurer is stationary in his laboratory or is traveling at high speed; or whether the source of light is stationary or moving with respect to the observer. Although the result may appear strange, it has been confirmed by many independent experiments.

Einstein combined the constancy of the speed of light in vacuum with a basic philosophical idea about the role of reference frames (discussed in Unit 1) in physical theory. He postulated that all reference frames that move with uniform velocity relative to each other are equivalent: no one of these frames is preferable to any other. This means that the laws of physics must be the same in all such reference frames. Another way of saying this is that the law of physics are invariant with respect to uniform motion, that is, they are not affected by uniform motion. It would be very inconvenient if this were not the case: for example, if Newton's laws of motion did not hold in a train moving at constant speed relative to the surface of the earth.

The combination of the idea of invariance with the constancy of the speed of light led Einstein to many remarkable

Summary 20.1
1. The special theory of relativity is based on two postulates. While the postulates sound quite simple, their consequences are most dramatic.

2. Mass, which is invariant in Newtonian mechanics, is predicted to increase with speed. Experimentally, it does increase precisely as predicted.

Most important to the story in this text is the equivalence of mass and energy (e.g. kinetic energy adds to an objects mass). The conversion factor from mass to energy is c^2.

The special theory of relativity, published in 1905, restricts itself to reference frames in uniform motion relative to each other. One of the main concerns of the special theory is electromagnetic phenomena. The special theory appeared after a period of intensive thought and investigation into the nature of electromagnetic wave propagation. One such investigation was the Michelson-Morley experiment mentioned in Unit 4.

The invariance of Maxwell's electromagnetic equations was one of the initial problems around which Einstein developed special relativity.

"Where you hide in the cellar and then look down
On the poets that live in the attics
For the whole of the house is upside down
In the Higher Mathematics."
G.K. Chesterton, 'Songs of Education'

Two postulates form the foundation of the special theory. From these two postulates, all the consequences (including those that seem bizaarre) can be logically deduced.
The first postulate is the constancy of the speed of light.
The second postulate maintains that the form of physical law is the same in all reference frames moving with a uniform velocity.

Over the years, different theoreticians have come up with alternative ways of establishing the postulates, but the flavor is the same.

The general theory of relativity, published in 1916, extends to reference frames in non-uniform (accelerated) motion. One of the main concerns of the general theory is gravitational phenomena. While there have been many experimental checks on the special theory, there have been only a few on the general theory.

results concerning our ideas of space and time, and to modifications of Newtonian mechanics. We cannot here go through the details of Einstein's work because too much time would be needed. We can, however, state some of the theoretical results he obtained and see if they agree with experiment. It is, after all, the comparison between theory and experiment which is a chief test of the relativity theory, as it is with any other theory in physics.

The most striking results of the relativity theory appear for bodies moving at very high speeds, that is, at speeds that are not negligible compared to the speed of light. For bodies moving at speeds small compared to the speed of light, relativity theory yields the same results as Newtonian mechanics as nearly as we can measure. This must be the case because we know that Newton's laws account very well for the motion of the bodies with which we are familiar in ordinary life. We shall, therefore, look for differences between relativistic mechanics and Newtonian mechanics in experiments involving high-speed particles. For the purposes of this course the differences are presented as deviations from classical physics and in the language of classical physics. Relativity involves, however, a large shift in viewpoint and in ways of talking about physics.

See "Mr Tompkins and Simultaneity" in Project Physics Reader 5.

We saw in Sec. 18.2 that J. J. Thomson devised a method for determining the speed v and the ratio of charge to mass q_e/m for electrons. Not long after the discovery of the electron by Thomson it was found that the value of q_e/m was not really constant, but varies with the speed of the electrons. Several physicists found, between 1900 and 1910, that electrons have the value $q_e/m = 1.76 \times 10^{11}$ coul/kg only for speeds that are very small compared to the speed of light; the ratio has smaller values for electrons with greater speeds. The relativity theory offered an explanation for these results. According to the theory of relativity, the electron charge does not depend on the speed of the electrons; but the mass of an electron should vary with speed, increasing according to the formula

See "Mathematics and Relativity" in Project Physics Reader 5.

$$m = \frac{m_o}{\sqrt{1 - v^2/c^2}} .$$

In this formula, v is the speed of the electron, c is the speed of light in vacuum and m_o is the rest mass, the electron mass when the electron is not moving, that is, when $v = 0$. More precisely, m_o is the mass of the electron when it is at rest with respect to an observer, to the person doing the experiment; m is the mass of the electron measured while it moves with speed v relative to the observer. We may call m the relativistic mass. It is the mass determined, for example, by means of J. J. Thomson's method.

SG 20.1

The ratio of relativistic mass to rest mass, m/m_o, which is equal to $1/\sqrt{1 - v^2/c^2}$, is listed in Table 20.1 for values of v/c which approach unity. The value of m/m_o becomes very large as v approaches c.

See "Relativity" in Project Physics Reader 5.

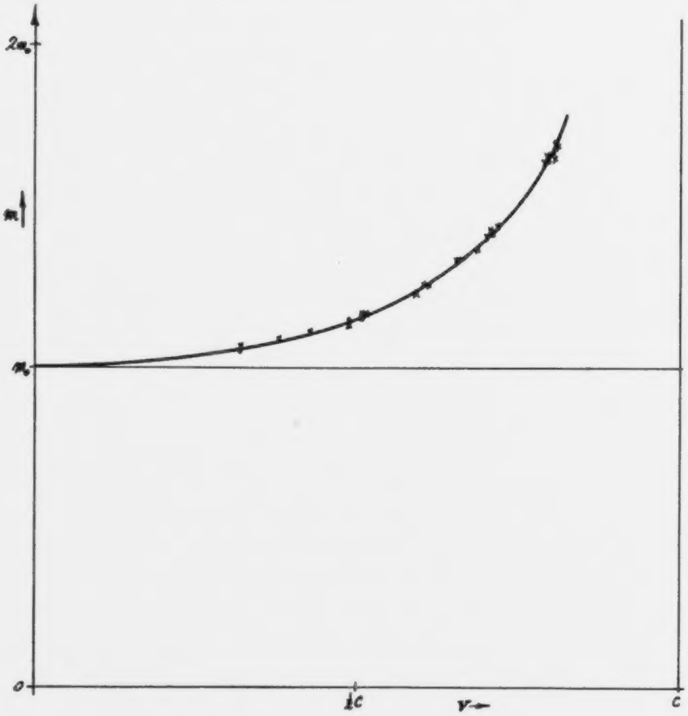
Table 20.1 The Relativistic Increase of Mass with Speed

v/c	m/m_o	v/c	m/m_o
0.0	1.000	0.95	3.203
0.01	1.000	0.98	5.025
0.10	1.005	0.99	7.089
0.50	1.155	0.998	15.82
0.75	1.538	0.999	22.37
0.80	1.667	0.9999	70.72
0.90	2.294	0.99999	223.6

See "Parable of the Surveyors" in Project Physics Reader 5.

See "Outside and Inside the Elevator" in Project Physics Reader 5.

The formula for the relativistic mass has been tested experimentally; some of the earlier results, for electrons with speeds so high that the value of v reaches about 0.8 c, are shown in the graph at the right. At that value of v the relativistic mass m is about 1.7 times the rest mass m_o. The curve shows the theoretical variation of m as the value of v increases, and the dots and crosses are results from two different experiments. The agreement of experiment and theory is excellent. The increase in mass with speed accounts for the shrinking of the ratio q_e/m with speed, which was mentioned earlier.

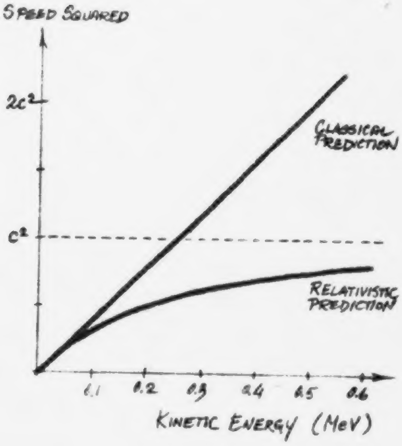


20.1

The theory of relativity says that the formula for variation of mass is valid for all moving bodies, not just electrons and other atomic particles. But larger bodies, such as those with which we are familiar in everyday life, move with speeds so small compared to that of light that the value of v/c is very small. The value of v^2/c^2 is then extremely small, and the values of m and m_o are so nearly the same that we cannot tell them apart. In other words, the relativistic increase in mass can be detected only for particles of sub-atomic size, which can move at very high speeds.

relative to us

The effects discussed so far are mainly of historical interest because they helped convince physicists of the correctness of relativity theory. Experiments done more recently provide even more striking evidence of the breakdown of Newtonian physics for particles with very high speeds. Electrons can be given very high energies by accelerating them by means of a high voltage V. Since the electron charge is known, the energy increase, q_eV, is known. The rest mass m_o of an electron is also known (see Sec. 18.3) and the speed v can be measured. It is, therefore, possible to compare the values of the energy q_eV with $\frac{1}{2}m_ov^2$. When experiments of this kind are done, it is found that when the electrons have speeds that are small compared to the speed of light, $\frac{1}{2}m_ov^2 = q_eV$. We used this relation in discussing the photoelectric effect. We could do so because photoelectrons do, indeed, have small speeds and m and m_o are very nearly identical for them. But, when the speed of the electron becomes large so that v/c is no longer small compared to 1.0, it is found that $\frac{1}{2}m_ov^2$ does not increase in proportion to q_eV; the discrepancy increases as q_eV increases. The increase in kinetic energy still equals the amount of electrical work done, q_eV, but some of the energy increase becomes measurable as the increase in mass instead of a marked increase in speed. The value of v^2, instead of steadily increasing with kinetic energy, approaches a limiting value: c^2.



In the Cambridge Electron Accelerator (CEA) operated in Cambridge, Massachusetts, by Harvard University and the Massachusetts Institute of Technology, electrons are accelerated in many steps to an energy which is equivalent to what they would gain in being accelerated by a potential difference of 6 × 10^9 volts—an enormous energy for electrons. (Unit 6 deals further with accelerators, and the operation of the CEA apparatus is also the subject of a movie "Synchrotron".) The speed attained by the electrons is

• But it is accelerated in many, many small steps

$v = 0.999999996\,c$; at this speed the relativistic mass m is over 10,000 times greater than the rest mass m_o!

SG 20.2

The energy equivalent of the electron's rest mass is about 500 KeV. The KE of an electron in a tv tube is about 20 KeV, so its mass is increased by about $\frac{20\text{ KeV}}{500\text{ KeV}}$, or 4%!

Relativity theory leads to a new formula for kinetic energy, expressing it in terms of the increase in mass:

$$KE = (m - m_o)c^2$$

or

$$KE = mc^2 - m_oc^2.$$

SG 20.3

It can be shown in a few steps of algebra that $mc^2 - m_oc^2$ is almost exactly equal to $\frac{1}{2}m_ov^2$ when v is very small compared to c. But at very high speeds, $mc^2 - m_oc^2$ agrees with experimental values of the amount of work done on a particle and $\frac{1}{2}m_ov^2$ does not. Einstein gave the following interpretation of the terms in the relativistic formula for KE: mc^2 is the total energy of the particle, and m_oc^2 is an energy the particle has even when it is at rest:

$$\begin{array}{ccccc} KE & = & mc^2 & - & m_oc^2 \\ \text{kinetic energy} & = & \text{total energy} & - & \text{rest energy} \end{array}$$

The rest energy m_0c^2 includes the potential energy, if there is any. Thus a compressed spring has a somewhat larger rest mass and rest energy than the same spring when relaxed.

Or, putting it the other way around, the total energy E of a particle is the sum of its rest energy and its kinetic energy:

$$\begin{aligned} E &= mc^2 \\ &= m_oc^2 + KE. \end{aligned}$$

This equation, Einstein's mass-energy relation, has great importance in nuclear physics. It suggests that kinetic energy can be converted into rest mass, and rest mass into kinetic energy or radiation. In Chapters 23 and 24, we shall see how such changes come about experimentally, and see additional experimental evidence which supports this relationship.

SG 20.4

SG 20.5

The theory of relativity was developed by Einstein from basic considerations of the nature of space and time and of their measurement. He showed that the Newtonian (or classical) views of these concepts led to contradictions and had to be revised. The formulas for the variation of mass with speed and the mass-energy relation resulted from the logical development of Einstein's basic considerations. The predictions of the theory have been verified experimentally, and the theory represents a model, or view of the world, which is an improvement over the Newtonian model.

SG 20.6

SG 20.7

Q1 What happens to the measurable mass of a particle as its kinetic energy is increased?

Q2 What happens to the speed of a particle as its kinetic energy is increased?

Summary 20.2
1. Since photons have energy, they ought also to have an equivalent mass and have momentum.

2. When a photon scatters from an electron, the interaction can be analyzed as a collision between two particles, using conservation of momentum and KE. (Note the wave-particle dualism: although the momentum of the photon is used in the calculation, it is its change in frequency which is calculated.)

20.2 Particle-like behavior of radiation. The first use we shall make of a result of relativity theory is in the further study of light quanta and of their interaction with atoms. The photoelectric effect taught us that a light quantum has energy hf, where h is Planck's constant and f is the frequency of the light. This concept also applies to x rays which, like visible light, are electromagnetic radiation, but of higher frequency. The photoelectric effect, however, didn't tell us anything about the momentum of a quantum. We may raise the question: if a light quantum has energy does it also have momentum?

The theory of relativity makes it possible for us to define the momentum of a photon. We start with the mass-energy relation for a particle, $E = mc^2$, and write it in the form:

$$m = \frac{E}{c^2}.$$

We may then speculate that the magnitude of the momentum p is

$$p = mv = \frac{E}{c^2} v.$$

Owing to the large value of "f" in the x-ray region, an x-ray photon may have a mass comparable to that of an electron at rest.

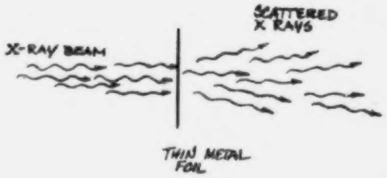
The last term is an expression for the momentum from which the mass has been eliminated. If this formula could be applied to a light quantum by setting the speed v equal to the speed of light c in the above equation; we would get

$$p = \frac{Ec}{c^2} = \frac{E}{c}.$$

Now, $E = hf$ for a light quantum, and if we substitute this expression for E in $p = E/c$, we would get for the momentum of a light quantum:

$$p = \frac{hf}{c}.$$

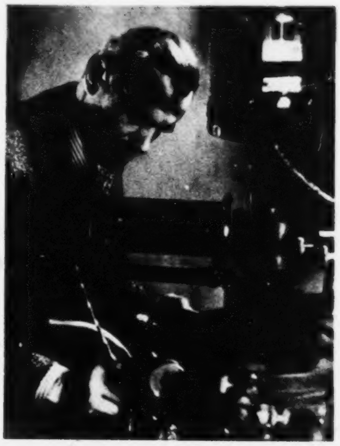
Does it make sense to define the momentum of a photon in this way? It does if the definition can be applied successfully to the interpretation of experimental results. The first example of the successful use of the definition was in the analysis of the Compton effect which will now be considered.



According to classical electromagnetic theory, when a beam of light (or x rays) strikes the atoms in a target (such as a thin sheet of metal), the light will be scattered in various directions but its frequency will not be changed. Light of a certain frequency may be absorbed by an atom, and light of another frequency may be emitted; but, if the light is simply scattered, there should be no change in frequency—provided that the classical wave theory is correct.

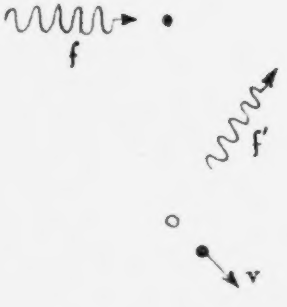
F44: Interference of photons

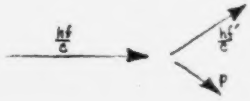
According to quantum theory, however, light is made up of photons. Compton reasoned that if photons have momentum, then in a collision between a photon and an atom the law of conservation of momentum should also apply. According to this law (see Chapter 10), when a body of small mass collides with a massive object, it simply bounces back or glances off with very little change in energy. But, if the masses of the two colliding objects are not very much different, a significant amount of energy can be transferred in the collision. Compton calculated how much energy a photon should lose in a collision with an atom, assuming that the energy and momentum of the photon are defined as hf and hf/c, respectively. The change in energy is too small to observe if a photon simply bounces off an entire atom. If, however, a photon strikes an electron, which has a small mass, the photon should transfer a significant amount of energy to the electron.



Arthur H. Compton (1892-1962) was born in Wooster, Ohio and graduated from the College of Wooster. After receiving his doctor's degree in physics from Princeton University in 1916, he taught physics and then worked in industry. In 1919-1920 he did research under Rutherford at the Cavendish Laboratory of the University of Cambridge. In 1923, while studying the scattering of x rays, he discovered and interpreted the changes in the wavelengths of x rays when the rays are scattered. He received the Nobel Prize in 1927 for this work.

In experiments up to 1923, no difference had been observed between the frequencies of the incident and scattered light (or x rays) when electromagnetic radiation was scattered by matter. In 1923 Compton, using improved experimental techniques, was able to show that when a beam of x rays of a given frequency is scattered, the scattered beam consists of two parts: one part has the same frequency as the incident x rays; the other part has slightly lower frequency. This reduction in frequency of some of the scattered x rays is called the Compton effect. The change of frequency corresponds to a transfer of energy from photons to electrons in accordance with the laws of conservation of momentum and energy. The observed change in frequency is just what would be predicted if the photons were particles having momentum $p = \frac{hf}{c}$. Furthermore, the electrons which were struck by the photons could also be detected, because they were knocked out of the target. Compton found that the momentum of these electrons was just what would be expected if they had been struck by a particle with momentum $p = \frac{hf}{c}$.





Compton's experiment showed that a photon can be regarded as a particle with a definite momentum as well as energy; it also showed that collisions between photons and electrons obey the laws of conversation of momentum and energy.

Photons act much like particles of matter, having momentum as well as energy; but they also act like waves, having frequency and wavelength. In other words, the behavior of electromagnetic radiation is sometimes similar

SG 20.8

The set of equations expressing conservation of energy and momentum can be solved to provide an expression for the wavelength shift: $\lambda' - \lambda = \frac{h}{mc}(1 - \cos\Theta)$ where θ is scattering angle
m is electron rest mass.

to what we are used to thinking of as particle behavior and sometimes similar to what we are used to thinking of as wave behavior. This behavior is often referred to as the wave-particle dualism of radiation. The question, "Is a photon a wave or a particle?" can only be answered: it may not *be* either, but can appear to *act* like either, depending on what we are doing with it.

SG 20.9

SG 20.10

Q3 How does the momentum of a photon depend on the frequency of the light?

Q4 What did the Compton effect prove?

Summary 20.3

1. De Broglie postulated that the wave-particle dualism also applied to particles; that is, a particle has a wavelength associated with it.

2. Diffraction patterns in beams of electrons gave experimental verification.

3 The quantized-orbits postulate of Bohr can be deduced (somewhat spuriously as it turns out) from the de Broglie postulate.

De Broglie's radical ideas were met with indifference: in fact, his dissertation was on the verge of being rejected. Fortunately, Einstein was asked to express an opinion on the ideas contained in the dissertation. Partly as a result of Einstein's enthusiastic response, de Broglie was granted the doctor's degree by the Sorbonne in 1924.

20.3 Wave-like behavior of matter. In 1924, a French physicist, Louis de Broglie, suggested that the wave-particle dualism which applies to radiation might also apply to electrons and other atomic particles. Perhaps, he said, this wave-particle dualism is a fundamental property of all quantum processes, and what we have always thought of as material particles sometimes act like waves. He then sought an expression for the wavelength of an electron and found one by means of a simple argument.

We start with the formula for the magnitude of the momentum of a photon,

$$p = \frac{hf}{c}.$$

The speed and frequency of a photon are related to the wavelength by the relation

$$c = f\lambda,$$

$$\text{or} \qquad \frac{f}{c} = \frac{1}{\lambda}$$

If we replace $\frac{f}{c}$ in the momentum equation by $\frac{1}{\lambda}$, we get:

$$p = \frac{h}{\lambda},$$

$$\text{or} \qquad \lambda = \frac{h}{p}.$$

The de Broglie wavelength of a material particle does not refer to light, but to some new wave property associated with the motion of matter itself.

De Broglie suggested that this relation, derived for photons, would also apply to electrons with the momentum $p = mv$. He, therefore, wrote for the wavelength of an electron:

$$\lambda = \frac{h}{mv},$$

where m is the mass of the electron and v its speed.

SG 20.11

What does it mean to say that an electron has a wavelength equal to Planck's constant divided by its momentum? If this statement is to have any physical meaning, it must

In case there is any doubt, the de Broglie waves for electrons, protons, neutrons, etc., are not electromagnetic waves.

Beams of neutrons whose wavelength is of the same order of magnitude as the spacing of atoms in a solid are conveniently obtained from a nuclear reactor. These beams show diffraction effects just like x rays do.

be possible to test it by some kind of experiment. Some wave property of the electron must be measured. The first such property to be measured was diffraction.

By 1920 it was known that crystals have a regular lattice structure; the distance between rows or planes of atoms in a crystal is about 10^{-10}m. After de Broglie proposed his hypothesis that electrons have wave properties, several physicists suggested that the existence of electron waves might be shown by using crystals as diffraction gratings. Experiments begun in 1923 by C. J. Davisson and L. H. Germer in the United States, yielded diffraction patterns similar to those obtained for x rays, as illustrated in the two drawings at the left below. The experiment showed not only that electrons do have wave properties, but also that their wavelengths are correctly given by de Broglie's relation, $\lambda = h/mv$. These results were confirmed in 1927 by G. P. Thomson, who directed an electron beam through thin gold foil to produce the more familiar type of diffraction pattern like the one at the right in the margin. By 1930, diffraction from crystals had been used to demonstrate the wave-like behavior of helium atoms and hydrogen molecules, as illustrated in the drawing at the right below.

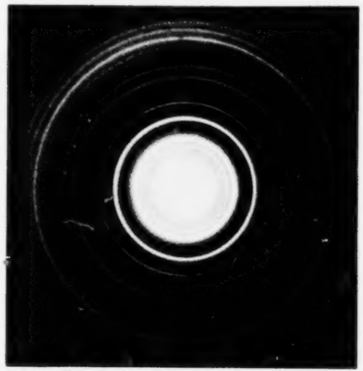
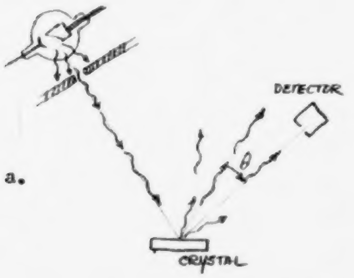


Fig. 20.3 Diffraction pattern produced by directing a beam of electrons through polycrystalline aluminum. With a similar pattern, G.P. Thomson demonstrated the wave properties of electrons—28 years after their particle properties were first demonstrated by J.J. Thomson, his father.

Although these setups would work, the actual experiments might have been considerably different.

a.



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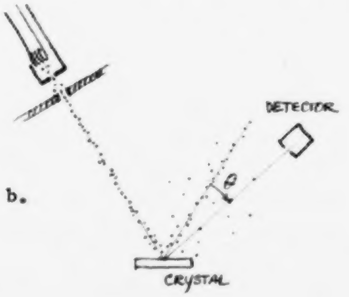
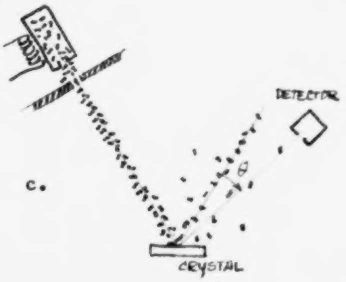


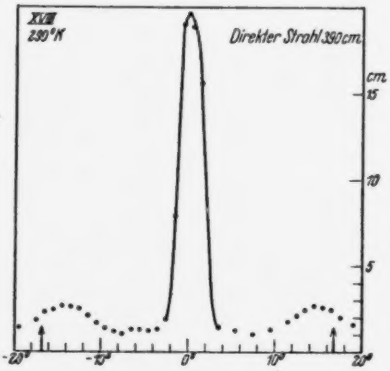
c.



a. One way to demonstrate the wave behavior of x rays is to direct a beam at the surface of a crystal. The reflections from different planes of atoms in the crystal interfere to produce reflected beams at angles other than the ordinary angle of reflection.

b. A very similar effect can be demonstrated for a beam of electrons. The electrons must be accelerated to an energy that corresponds to a deBroglie wavelength of about 10^{-10} m (which requires an accelerating voltage of only about 100 volts).

c. More surprisingly still, a beam of molecules directed at a crystal will show a similar diffraction pattern. The diagram above shows how a beam of hydrogen molecules (H_2) can be formed by slits at the opening of a heated chamber; the average energy of the molecules is controlled by adjusting the temperature of the oven. The graph, reproduced from Zeitschrift für Physik, 1930, shows results obtained by I. Estermann and O. Stern in Germany. The detector reading is plotted against the deviation to either side of angle of ordinary reflection.



Diffraction pattern for H_2 molecules glacing off a crystal of lithium fluoride.

An electron which has a kinetic energy of 100 eV has a momentum $mv = 9.1 \times 10^{-31} \times 5.9 \times 10^{6} = 5.4 \times 10^{-24}$ Kg.m/sec.

hence $\lambda = \frac{h}{mv} = \frac{6.6 \times 10^{-34}}{5.4 \times 10^{-24}} = 1.2 \times 10^{-10} m = 1.2\ \text{Å}$

This is comparable to the wavelength of x rays.

V was obtained from the non-relativistic expansion

$$V = \sqrt{\frac{2E}{m}}$$

The de Broglie wavelength: examples.

A body of mass 1 kg moves with a speed of 1 m/sec. What is its de Broglie wavelength?

$$\lambda = \frac{h}{mv}$$

$$h = 6.6 \times 10^{-34} \text{ joule·sec}$$

$$mv = 1 \text{ kg·m/sec}$$

$$\lambda = \frac{6.6 \times 10^{-34} \text{ joule·sec}}{1 \text{ kg·m/sec}},$$

or

$$\lambda = 6.6 \times 10^{-34} \text{ m.}$$

The de Broglie wavelength is much too small to be detected. We would expect to detect no wave aspects in the motion of this body.

An electron of mass 9.1×10^{-31} kg moves with a speed of 2×10^{6} m/sec. What is its de Broglie wavelength?

$$\lambda = \frac{h}{mv}$$

$$h = 6.6 \times 10^{-34} \text{ joule·sec}$$

$$mv = 1.82 \times 10^{-24} \text{ kg·m/sec}$$

$$\lambda = \frac{6.6 \times 10^{-34} \text{ joule·sec}}{1.82 \times 10^{-24} \text{ kg·m/sec}},$$

or

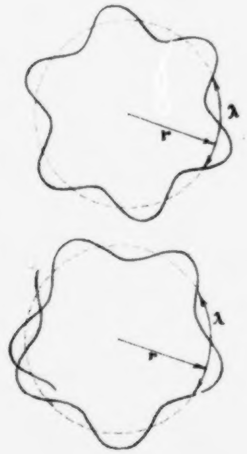
$$\lambda = 3.6 \times 10^{-10} \text{ m.}$$

The de Broglie wavelength is of atomic dimensions; for example, it is of the same order of magnitude as the distances between atoms in a crystal. We would expect to see wave aspects in the interaction of electrons with crystals.

SG 20.12
SG 20.13

According to de Broglie's hypothesis, which has been confirmed by these experiments, wave-particle dualism is a general property not only of radiation but also of matter. It is now customary to use the word "particle" to refer to electrons and photons while recognizing that they both have properties of waves as well as of particles.

Fig. 20.4 Only certain wavelengths will "fit" around a circle.



De Broglie's relation, $\lambda = \frac{h}{mv}$, has an interesting yet simple application which makes more reasonable Bohr's postulate that the angular momentum of the electron in the hydrogen atom can only have certain values. Bohr assumed that the angular momentum can have only the values:

$$mvr = n\frac{h}{2\pi}, \text{ where } n = 1, 2, 3, \ldots$$

Now, suppose that an electron wave is somehow spread over an orbit of radius r—that, in some sense, it "occupies" an orbit of radius r. We may ask if standing waves can be set up as indicated, for example, in Fig. 20.4. The condition for such standing waves is that the circumference of the orbit is equal in length to a whole number of wavelengths, that is, to nλ. The mathematical expression for this condition is:

$$2\pi r = n\lambda.$$

When this wave comes back to its starting point, crest and trough do not exactly come together. The waves tend to cancel and no fixed pattern results. This concept, like the Bohr orbit concept, is now regarded as oversimplified, as you will see in Sec 20.6. The nodes of the standing electron wave are spherical surfaces or planes. The standing wave on the coffee cup, which appear on the cover of the Unit 3 Student Handbook, is a closer analogy.

F4.5: Matter waves

20.4

If we now replace λ by $\frac{h}{mv}$, according to de Broglie's relation, we get

$$2\pi r = n\frac{h}{mv},$$

or

$$mvr = n\frac{h}{2\pi}.$$

But, this is just Bohr's quantization condition! The de Broglie relation for electron waves allows us to derive the quantization that Bohr had to assume.

The result obtained indicates that we may picture the electron in the hydrogen atom in two ways: either as a particle moving in an orbit with a certain angular momentum, or as a standing de Broglie type wave occupying a certain region around the nucleus.

Q5 Where did de Broglie get the relation $\lambda = \frac{h}{mv}$ for electrons?

Q6 Why were crystals used to get diffraction patterns of electrons?

20.4 Quantum mechanics. The proof that things (electrons, atoms, molecules) which had been regarded as particles also show properties of waves has served as the basis for the currently accepted theory of atomic structure. This theory, quantum mechanics, was introduced in 1925; it was developed with great rapidity during the next few years, primarily by Heisenberg, Born, Schrödinger, Bohr and Dirac. The theory appeared in two different mathematical forms proposed independently by Heisenberg and Schrödinger. These two forms were shown by Dirac to be equivalent. The form of the theory that is closer to the ideas of de Broglie, discussed in the last section, was that of Schrödinger. It is often referred to as "wave mechanics."

Summary 20.4

1. The theory that correlates both the wave and particle nature of matter is quantum mechanics.

2. Schrödinger's form of the theory, which is closely related to the de Broglie hypothesis, successfully predicts the results of the Bohr theory for hydrogen; however, it does much more.

3. Quantum mechanics provides the present framework for our understanding of atomic structure.

Schrödinger sought to express the dual wave and particle nature of matter mathematically by means of a wave equation. Maxwell had formulated the electromagnetic theory of light in terms of a wave equation, and physicists were familiar with this theory and its applications. Schrödinger reasoned that a wave equation for electrons would have to resemble the wave equation for light, but would have to include Planck's constant to permit quantum effects. Now, the equations we are talking about are not algebraic equations. They involve higher mathematics and are called "differential equations." We cannot discuss this mathematical part of wave mechanics, but the physical ideas involved require only a little mathematics and are essential to an understanding of modern physics. So, in the rest of

Heisenberg developed an alternate approach to quantum mechanics. His approach grew more directly from the experimental data of atomic spectra. The two approaches, Schrödinger's and Heisenberg's, were ultimately shown by Dirac to be equivalent.

D58: Ionization potential

Paul Adrien Maurice Dirac (1902-), an English physicist, was one of the developers of modern quantum mechanics. His relativistic theory of quantum mechanics (1930) was the first indication that "anti-particles" exist, such as the positron. He shared the Nobel Prize for physics in 1933 with Schrödinger. In 1932, at the age of 30, Dirac was appointed Lucasian Professor of Mathematics at Cambridge University, the post held by Newton.

Max Born (1882-) was born in Germany, but left that country in 1933 when Hitler and the Nazis gained control. Born was largely responsible for introducing the statistical interpretation of wave mechanics. From 1933 to 1953, when he retired, he worked at Cambridge, England and Edinburgh, Scotland. He was awarded the Nobel Prize in physics in 1954.

Prince Louis Victor de Broglie (1892-) comes of a noble French family. His ancestors served the French kings as far back as the times of Louis XIV. He was educated at the Sorbonne in Paris, served as a radio specialist in World War I, and was awarded the Nobel Prize in physics in 1929.

Werner Karl Heisenberg (1901 -), a German physicist, was one of the developers of modern quantum mechanics (at the age of 23). He discovered the uncertainty principle, and after the discovery of the neutron in 1932, proposed the proton-neutron theory of nuclear constitution. He was awarded the Nobel Prize in physics in 1932. During World War II, Heisenberg was in charge of German research of the application of nuclear energy.

Erwin Schrödinger (1887-1961) was born in Austria. After service in World War I, he became a professor of physics in Germany. He developed wave mechanics in 1926, left Germany in 1933 when Hitler and the Nazis came to power. From 1940 to 1956, when he retired, he was professor of physics at the Dublin Institute for Advanced Studies. He shared the Nobel Prize in physics with Dirac in 1933 for his work on wave mechanics.

this chapter, we shall discuss some of the physical ideas of the theory to try to make them seem plausible; and we shall consider some of the results of the theory and some of the implications of these results.

F46: Light: wave and quantum theories

Schrödinger was successful in deriving an equation for the motions of electrons. This equation, which has been named after him, defines the wave properties of electrons and also includes their basic particle aspects. The mathematical solution of the Schrödinger equation shows that only certain electron energies are possible in an atom. For example, in the hydrogen atom, the single electron can only be in those states for which the energy of the electron has the values:

$$E_n = - \frac{2\pi^2 m q_e{}^4}{n^2 h^2} ,$$

with n having only whole number values. These values of the energies are just the ones given by the Bohr theory. But, in Schrödinger's theory, this result follows directly from the mathematical formulation of the wave and particle nature of the electron. The existence of these stationary states has not been assumed, and no assumptions have been made about orbits. The new theory yields all the positive results of the Bohr theory without having any of the inconsistent hypotheses of the earlier theory. The new theory also accounts for the experimental information for which the Bohr theory failed to account.

See "The New Landscape of Science" in Project Physics Reader 5.

On the other hand, quantum mechanics does not supply a physical model or picture of what is going on inside the atom. The planetary model of the atom has had to be given up, and has not been replaced by another simple picture. There is now a highly successful mathematical model, but no easily understood physical model. The concepts used to build quantum mechanics are more abstract than those of the Bohr theory; it is hard to get an intuitive feeling for atomic structure. But the mathematical theory of quantum mechanics is much more powerful than the Bohr theory, and many problems have been solved with quantum mechanics that were previously unsolvable. Physicists have learned that the world of atoms, electrons and photons cannot be thought of in the same mechanical terms as the world of everyday experience. In fact the world of atoms has presented us with some new and fascinating concepts which will be discussed in the next two sections.

Q7 The set of energy states of hydrogen could be derived from Bohr's postulate of quantized angular momentum. Why was the derivation from Schrödinger's equation so much better?

Q8 Quantum (or wave) mechanics has had great success. What is its major drawback?

"... which rightly gazed upon
Show nothing but confusion,
Eyed awry,
Distinguished form, ..."
William Shakespeare,
'King Richard the Second'

20.5 Quantum mechanics - the uncertainty principle. The success of wave mechanics emphasizes the fundamental importance of the dual wave-and-particle nature of radiation and matter. The question now arises of how a particle can be thought of as "really" having wave properties. The answer is that invisible matter of the kinds involved in atomic structure doesn't have to be thought of as "really" being either particles or waves. Our ideas of waves and particles are taken from the world of visible things and may just not apply on the atomic scale. The suitability of applying wave and particle concepts to atomic problems has to be studied and its possible limitations determined.

Summary 20.5
The wave-particle dualism is an example of the difficulties encountered when we try to project our common sense notions, developed from the familiar, into the unknown.

See "Dirac and Born" in Project Physics Reader 5.

When we try to describe something that no one has ever seen or can ever see directly, it is questionable whether the concepts of the visible world can be taken over unchanged. It appeared natural before 1925 to try to talk about the transfer of energy in either wave terms or particle terms, because that was all physicists knew and understood at the time. No one was prepared to find that both wave and particle descriptions could apply to light and to matter. But this dualism cannot be wished away, because it is based on experimental results.

Max Born, one of the founders of quantum mechanics, has written: "The ultimate origin of the difficulty lies in the fact (or philosophical principle) that we are compelled to use the words of common language when we wish to describe a phenomenon, not by logical or mathematical analysis, but by a picture appealing to the imagination. Common language has grown by everyday experience and can never surpass these limits. Classical physics has restricted itself to the use of concepts of this kind; by analyzing visible motions it has developed two ways of representing them by elementary processes: moving particles and waves. There is no other way of giving a pictorial description of motions—we have to apply it even in the region of atomic processes, where classical physics breaks down."

If we didn't feel uncomfortable with the dualism, we could just accept it as a fact of nature and go on from there. But, scientists were as uncomfortable with the dualism as you undoubtedly are, and searched for a way out of the situation. Because there is no argument with the facts, the way out has to be with our view of nature, our outlook as scientists. To look for this way we shall describe some experiments which show up a fundamental limitation on our ability to describe phenomena. Following that, we shall give a simplified version of the present view of physics concerning the wave-particle dualism.

Up to this point we have always talked as if we could measure any physical property as accurately as we pleased, if not in the laboratory at least in our own thoughts in which ideal instruments could be "used." Wave mechanics shows that, even in thought experiments, there are limitations on the accuracy with which atomic measurements may be made.

20.5

Suppose we want to measure the position and velocity of a car; and let us suppose that the car's position is to be measured from the end of a garage. The car moves slowly out of the garage along the driveway. We mark the position of the front end of the car at a given instant by making a scratch on the ground; at the same time, we start a stopwatch. Then we run to the far end of the driveway, and at the instant that the front end of the car reaches another mark on the ground we stop the watch. We then measure the distance between the marks and get the average speed of the car by dividing the distance traversed by the time elapsed. Since we know the direction of the car's motion, we know the average velocity. Thus we know that at the moment the car reached the second mark it was at a known distance from its starting point and had traveled at a known *average velocity*.

How did we get this information? We could locate the position of the front end of the car because sunlight bounced off the front end into our eyes and permitted us to see when the car reached a certain mark on the ground. To get the average speed we had to locate the front end twice.

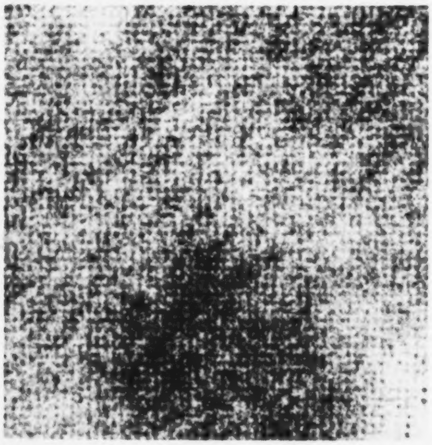
Note that we used sunlight in our experiment. Suppose that we had decided to use radio waves instead of light of visible wavelength. At 1000 kilocycles per second, a typical value for radio signals, the wave length is 300 meters. With radiation of this wavelength, which is very much greater than the dimensions of the car, it is impossible to locate the car with any accuracy, because the wavelength has to be comparable with or smaller than the dimensions of the object before the object can be located. *Radar* uses wavelengths from 3 cm to about 0.1 cm. Hence a radar apparatus could have been used instead of sunlight, but radar waves much longer than 3 cm would result in appreciable uncertainty about the positions and average speed of the car.

$$\lambda = \frac{c}{f} = \frac{3 \times 10^8 \text{ m/sec}}{10^6/\text{sec}} = 300 \text{ m.}$$

Let us now replace the car, driveway and garage by an electron leaving an electron gun and moving across an evacuated tube. We try to measure the position and speed of the electron. But some changes have to be made in the method of measurement. The electron is so small that we cannot locate its position by using visible light. The reason is that the wavelength of visible light is at least 10^4 times greater than the diameter of an atom.

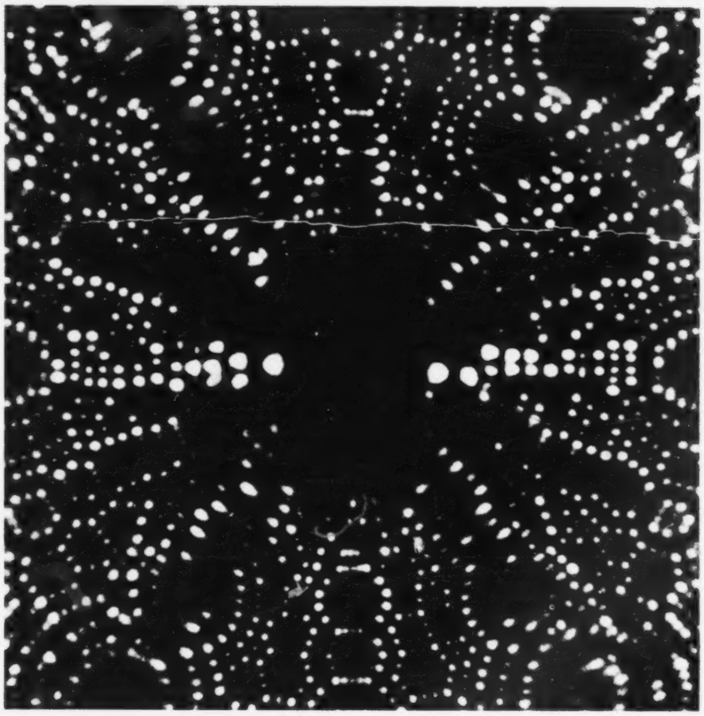
To locate an electron within a region the size of an atom (10^{-10} m) we must use a light beam whose wavelength is comparable to the size of the atom, if not much smaller. Otherwise we will be uncertain about the position by an

The extreme smallness of the atomic scale is indicated by these pictures made with techniques that give the very limits of magnification—about 10,000,000 times in this reproduction.



Electron micrograph of a section of a single gold crystal. The entire section of crystal shown is only 100Å across—smaller than the shortest wavelength of ultraviolet light that could be used in a light microscope. The finest detail that can be resolved is just under 2Å, so that the layers of gold atoms (spaced slightly more than 2Å) show as a checked pattern; individual atoms are beyond the resolving power.

This picture was taken with a field-ion microscope, in which a metallic tip is raised to a very high positive potential. The tip is housed in a vacuum chamber into which helium is introduced to a pressure of 10^{-3} mm of Hg. Helium ions are repelled from the region of strongest field on the tip, and travel radially out to a fluorescent screen. The high-field regions are the corners of the crystal array.



Field-ion micrograph of the tip of a microscopically thin tungsten crystal. As above, the entire section shown is only about 100Å across. The bright spots indicate the locations of atoms along edges of the crystal, but should not be thought of as pictures of the atom.

amount many times greater than the diameter of the electron. Now a photon of such a wavelength has a very great energy and momentum; and, from our study of the Compton effect, we know that the photon will give the electron a strong kick. SG 20.14 As a result, the velocity of the electron will be seriously changed, and in an unknown direction. Hence, although we have "located" the electron, we have altered its velocity (both in magnitude and direction). To say this more directly: the more accurately we locate the electron (by using photons of shorter wavelength) the less accurately we can know its velocity. If we try to disturb the electron less by using less energetic (longer wavelength) photons, we lose resolving power and acquire a greater uncertainty in the position of the electron. To summarize: we are unable to measure both the position and velocity of an electron to a prescribed accuracy. This conclusion is known as the uncertainty principle, and was discovered by Heisenberg. The uncertainty principle can be expressed quantitatively in a simple formula. If Δx is the uncertainty in position, and Δp is the uncertainty in momentum, then the product of the two must be equal to, or greater than, Planck's constant divided by 2π:

$$(\Delta x)(\Delta p) \geq \frac{h}{2\pi}$$

SG 20.15
SG 20.16
SG 20.17

The same reasoning holds for the car, but the limitation is of no practical consequence with such a massive object. It is only in the atomic world that the limitation is important.

Q9 If light photons used in finding the velocity of an electron disturb the electron too much, why can't the observation be improved by using weaker photons?

Q10 If the wavelength of light used to locate a particle is too long, why can't the location be found more precisely by using light of shorter wavelength?

Summary 20.6
Quantum mechanics provides the procedures for calculating the probability of electrons behaving in specific ways; thus with large numbers of electrons, their average behavior can be accurately described.

20.6 Quantum mechanics - probability interpretation. The way in which physicists now think about the dualism involves the idea of probability. Even in situations in which no single event can be predicted with certainty, it may still be possible to make predictions of the statistical probabilities of certain events. For example, automobile manufacturers don't know which 8 million people will buy cars this year. But they do know that about that many people will find need for a new car. Similarly, on a holiday weekend during which perhaps 25 million cars are on the road, the statisticians report a high probability that about 600 people will be killed in accidents. It isn't known which cars in which of

There is no uncertainty relation between any pair of variables, but only between certain pairs. Besides position and momentum, there are uncertainty relations, for instance, between energy and time, angular momentum and angle: $(\Delta E)(\Delta t) \geq \frac{h}{2\pi}$ $(\Delta L)(\Delta \theta) \geq \frac{h}{2\pi}$.

The uncertainty principle: examples.

A large mass.

Consider a car, with a mass of 1000 kg, moving with a speed of about 1 m/sec. Suppose that the uncertainty Δv in the speed is 0.1 m/sec (10% of the speed). What is the uncertainty in the position of the car?

$$\Delta x \Delta p \geq \frac{h}{2\pi}$$

$$\Delta p = m\Delta v = 100 \text{ kg·m/sec}$$

$$h = 6.63 \times 10^{-34} \text{ joule·sec}$$

$$\Delta x \geq \frac{6.63}{6.28} \times \frac{10^{-34} \text{ joule·sec}}{10^{2} \text{ kg·m/sec}}$$

or

$$\Delta x \geq 1 \times 10^{-36} \text{ m.}$$

The uncertainty in position is much too small to be observable. In this case we can determine the position of the body with as high an accuracy as we would ever need.

A small mass.

Consider an electron, with a mass of 9.1×10^{-31} kg, moving with a speed of about 2×10^{6} m/sec. Suppose that the uncertainty Δv in the speed is 0.2×10^{6} m/sec (10% of the speed). What is the uncertainty in the position of the electron?

$$\Delta x \Delta p \geq \frac{h}{2\pi}$$

$$\Delta p = m\Delta v = 1.82 \times 10^{-25} \text{ kg·m/sec}$$

$$h = 6.63 \times 10^{-34} \text{ joule·sec}$$

$$\Delta x \geq \frac{6.63}{6.28} \times \frac{10^{-34} \text{ joule·sec}}{1.82 \times 10^{-25} \text{ kg·m/sec}},$$

or

$$\Delta x \geq 5 \times 10^{-10} \text{ m.}$$

The uncertainty in position is of the order of atomic dimensions, and is significant in atomic problems.

The reason for the difference between these two results is that Planck's constant h is very small: so small that the uncertainty principle becomes important only on the atomic scale.

The main use of the uncertainty principle is in general arguments in atomic theory rather than in particular numerical problems. We don't really need to know exactly where an electron is, but we sometimes want to know if it could be in some region of space.

the 50 states will be the ones involved in the accidents, but the average behavior is still quite accurately predictable.

It is in this way that physicists think about the behavior of photons and material particles. As we have seen, there are fundamental limitations on our ability to describe the behavior of an individual particle. But the laws of physics often enable us to describe the behavior of large collections

See "The Fundamental Idea of Wave Mechanics" in Project Physics Reader 5.

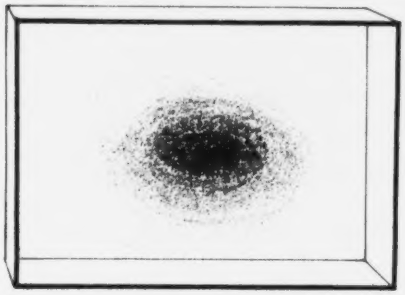
The quantum effect of discrete energy levels will occur in theory for any confined particle; but for particles large enough to be observed in a microscope, the quantum energy differences are so small that it would be impossible to observe even a lowest energy state or the gaps between energy states.

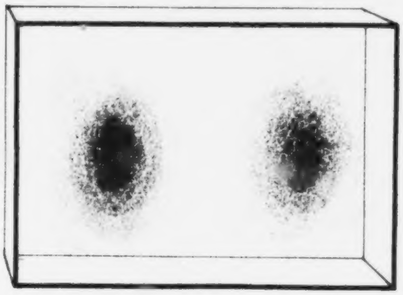
Probability in Quantum Mechanics

We have already described how probabilities were used in the kinetic theory of gases (Chapter 11). Because a gas contains so many molecules—more than a billion billion in each cubic centimeter of the air we breathe—it is impractical to calculate the motion of each molecule. Instead of applying Newton's laws to trace the paths of these molecules the scientists who developed the kinetic theory assumed that the net effect of all of the collisions among molecules would be a random, disordered motion that could be treated statistically. In the kinetic theory a gas is described by stating its average density and average kinetic energy, or, where more detail is wanted, by showing the relative numbers of molecules with different speeds.

Probability is used in a different way in quantum theory. The description of a single molecule or a single electron is given in terms that yield only statistical predictions. Thus quantum mechanics predicts the probability of finding a single electron in a given region. The theory does not specify the position and the velocity of the electron, but the probability of its having certain positions and certain velocities. The theory asserts that to ask for the precise position and velocity of a particle is to demand the unknowable.

As an example, consider the case of a particle confined to a box with rigid sides. According to classical mechanics the path of the particle can be traced from a knowledge of its position and velocity at some instant. Only if we introduce a large number of particles into the box is there a need to use probabilities.



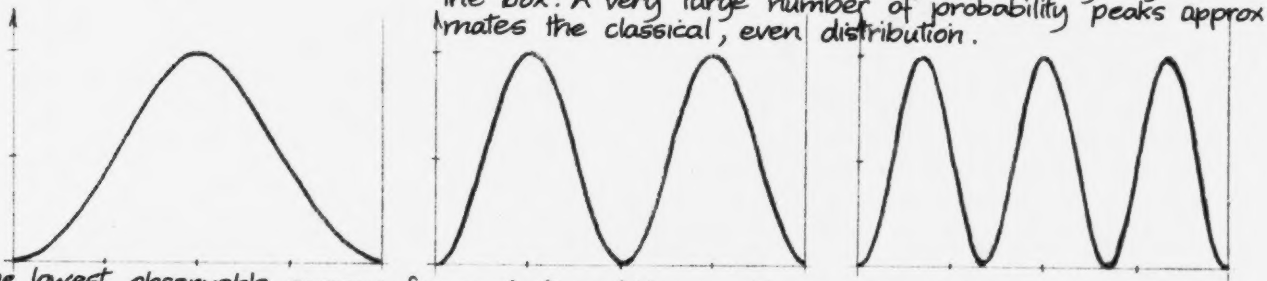


The quantum mechanical treatment of a single particle confined to a box is much different. It is not possible, according to the theory, to describe the particle as moving from one point to another within the box; only the probability of detecting the particle at various regions can be predicted. Moreover, the theory indicates that the particle is limited to certain discrete values of kinetic energy. The way the probability of finding the particle varies from point to point within the box depends on the energy. For example, in the lowest possible energy state the particle has the probability distribution indicated by the shading in the top drawing at the left; the darker the shading, the greater the probability of the particle's being there. The probability falls to zero at the sides of the box. The lower drawing at the left represents the probability distribution for the second energy level; notice that the probability is zero also for the particle to be on the center line.

As these drawings suggest, the probability distributions are the same as the intensity of standing waves that have nodes on the faces of the box. The standing wave intensity patterns for three of the lower energy levels are graphed below. The momentum and kinetic energy of the electron are connected to the wavelength of the standing waves through the deBroglie relations: $p = h/\lambda$ and $KE = h^2/2m\lambda^2$. Since only certain wavelengths can be fitted into the box, the particle can have only certain values of momentum and energy.

This quantum effect of discrete energy levels will occur, in theory, for any confined particle. Yet for a particle large enough to be seen with a microscope there does not appear to be any lower limit to its energy or any gaps in possible values of its energy. This is because the energy for the lowest state of such a particle is immeasurably small and the separation of measurably larger energies is also immeasurably small. The existence of discrete energy states can be demonstrated experimentally only for particles of very small mass confined to very small regions—that is, particles on the atomic scale.

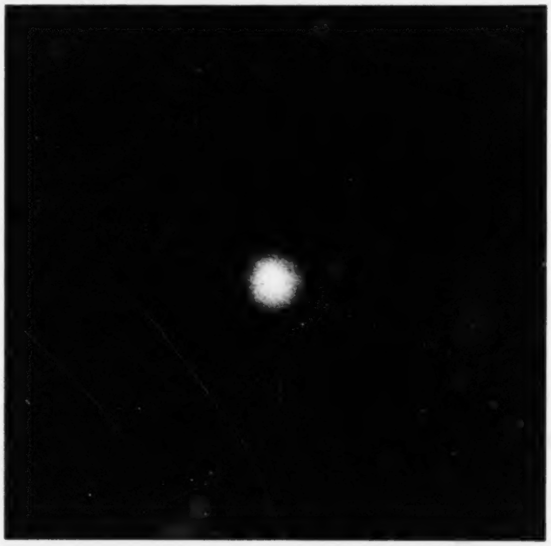
Classical picture gives equal probabilities of being anywhere in the box. A very large number of probability peaks approximates the classical, even distribution.



The lowest observable energy for a dust particle would correspond to, say, the trillionth energy state, for which a trillion standing wave peaks would be compressed into the box and the addition of another peak, increasing the energy to the next allowed value, would make no measurable difference.

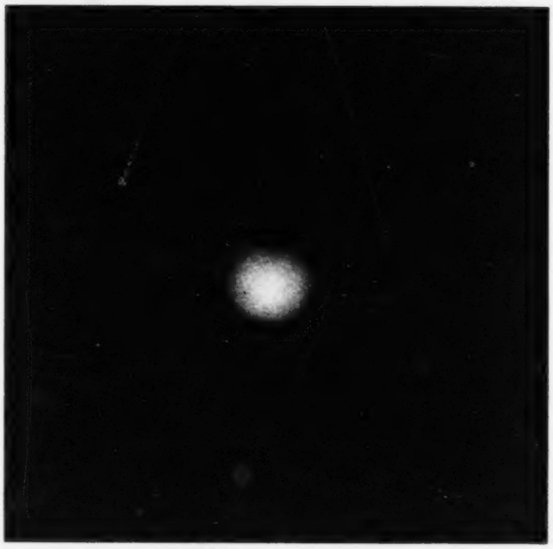
The two pictures represent s-states ($l = 0$), ie., those states which have spherically symmetrical charge distributions. Such a state with principal quantum number n can have up to $n-1$ nodes in its charge distribution.

Nowhere is the discreteness of energy states more pronounced than for electrons bound in atoms. The electron mass is extremely small and an atom makes an extremely small "box." There is thus clearly a lower limit to the energy of an electron in an atom and there are distinct gaps between energy levels.

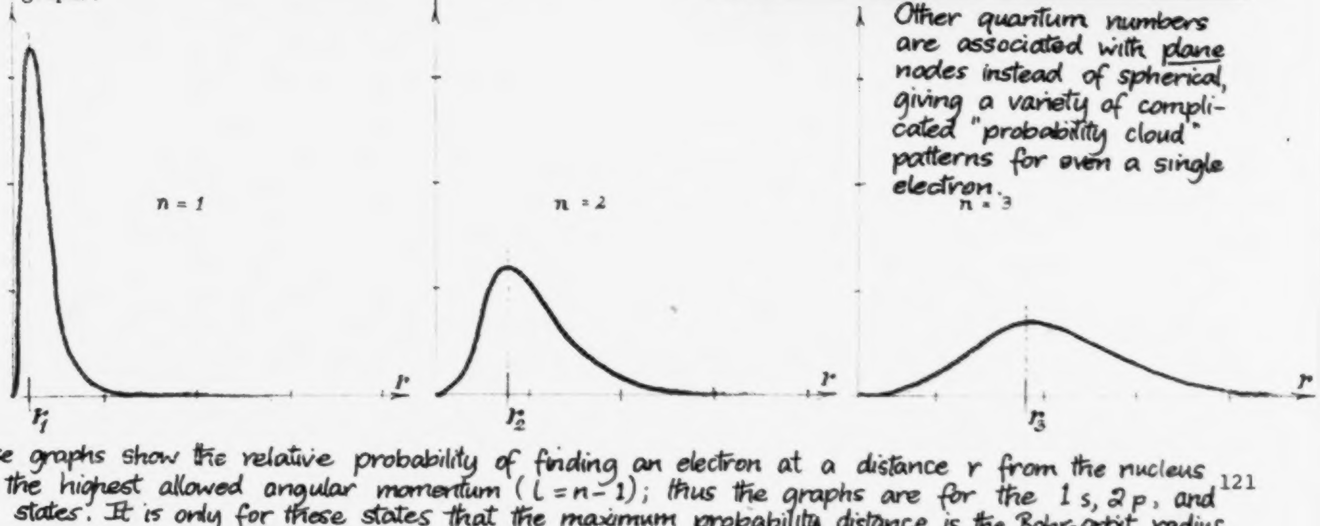


According to modern quantum theory, the hydrogen atom does not consist of a localized negative particle moving around a nucleus as in the Bohr model. Indeed, the theory does not provide any picture of the hydrogen atom. However, quantum theory does yield probability distributions similar to those on the preceding page. A description of this probability distribution is the closest thing that the theory provides to a picture. The probability distribution for the lowest energy state of the hydrogen atom is represented in the upper drawing at the right, where whiter shading at a point indicates greater probability. The probability distribution for a higher energy state, still for a single electron, is represented in the lower drawing at the right.

Quantum theory is, however, not really concerned with the position of any individual electron in any individual atom. Instead, the theory gives a mathematical representation that can be used to predict interaction with particles, fields and radiation. For example, it can be used to calculate the probability that hydrogen will emit light of a particular wavelength; the intensity and wavelength of light emitted by a large number of hydrogen atoms can then be compared with these calculations. Comparisons such as these have shown that the theory agrees with experiment.



Although the atom of modern quantum mechanics differs fundamentally from the Bohr model, there are points of correspondence between the two theories. The probability of finding the electron somewhere on a sphere at a distance r from the nucleus is plotted for the lowest energy state of the hydrogen atom at the left below. The most probable distance (r_1) is equal to the radius of the electron orbit given by the Bohr theory. The same correspondence occurs for higher energy states, as shown in the other two graphs.



Other quantum numbers are associated with plane nodes instead of spherical, giving a variety of complicated "probability cloud" patterns for even a single electron.

These graphs show the relative probability of finding an electron at a distance r from the nucleus for the highest allowed angular momentum ($l = n-1$); thus the graphs are for the 1s, 2p, and 3d states. It is only for these states that the maximum probability distance is the Bohr-orbit radius.

of particles with high accuracy. The solutions of the Schrödinger equation give us the probabilities for finding the particles at a given place at a given time.

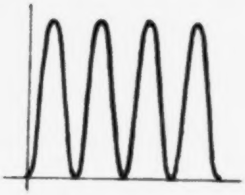
SG 20.18

To see how probability enters the picture we shall first examine a well-known problem from the point of view of waves. Then we shall examine the same situation from the point of view of particles.

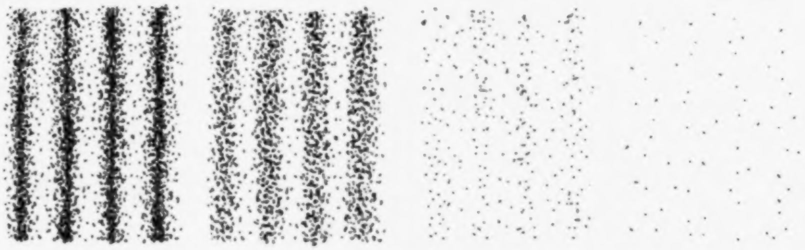
Strictly speaking, the probability is always zero for being exactly at any position. The graphs on these pages are of probability density, say D. The probability P of finding a particle in a small region of space Δs is $P = D \cdot \Delta s$. This idea is probably too subtle for many students, but the flavor can be adequately caught by referring to relative probability.

Imagine a television screen with a stream of electrons scanning it. The electron waves from the gun cover the screen with varying intensities to make the picture pattern. If the overall intensity of the waves is reduced by reducing the flow of electrons from the gun, the wave theory predicts that the picture pattern will remain, but that the entire picture will be fainter. If we were actually to do this experiment, we would find that, as the intensity becomes very weak, the picture pattern fades into a collection of separate faint flashes scattered over the screen. The naked eye is not sensitive enough to see the scattered flashes.

The waves give us the probability of finding electrons at various places at various times. If the number of electrons is small, the prediction becomes very poor. We can predict with any accuracy only the average behavior of large numbers of electrons.



This graph for a pattern of stripes would be interpreted in a wave model as the relative wave intensity, and in a particle model as the relative probability of a particle arriving.



A similar analysis holds for photons and their associated light waves. If light waves are projected onto a movie screen, the pattern is similar for all light intensities which give large numbers of photons. If the projector's light bulb is screened or otherwise reduced in intensity so that the light is extremely weak, the pattern falls apart into a collection of flashes. Here, too, the wave gives the probability of finding photons at various places at various times, and this probability gives us the correct pattern for large numbers of photons.

If a camera were pointed at the screen and the shutter left open for long enough time so that many photons (or electrons, in the previous example) arrived at the screen, the resultant picture would be a faithful reproduction of the high intensity picture. Even though individual particles arrive at random places on the screen, the rate at which they arrive doesn't affect the final result provided that we wait until the number that has finally arrived is very large. SG 20.19

We see then that we can deal only with the average behavior of atomic particles; the laws governing this average behavior turn out to be those of wave mechanics. The waves, it seems, are waves of probability. The probability that a particle will have some position at a given time travels through space in waves which interfere with each other in exactly the same way that water waves do. So, for example, if we think of electron paths crossing each other, we consider the electrons to be waves and compute the interference patterns which determine the directions in which the waves will be going after they have passed each other. Then, as long as there is no more interaction of the waves with matter, we can return to our description in terms of particles and say that the electrons end up going in such and such directions with such and such speeds. SG 20.20

We quote Max Born who was the originator of the probability interpretation of the wave-particle dualism:

> Every process can be interpreted either in terms of corpuscles or in terms of waves, but...it is beyond our power to produce proof that it is actually corpuscles or waves with which we are dealing, for we cannot simultaneously determine all the other properties which are distinctive of a corpuscle or of a wave, as the case may be. We can, therefore, say that the wave and corpuscular descriptions are only to be regarded as complementary ways of viewing one and the same objective process....

SG 20.21
SG 20.22

Despite the successes of the idea that the wave represents the probability of finding its associated particle in some specific condition of motion, many scientists found it hard to accept the idea that men cannot know exactly what any one particle is doing. The most prominent of such disbelievers was Einstein. In a letter to Born written in 1926, he remarked,

> The quantum mechanics is very imposing. But an inner voice tells me that it is still not the final truth. The theory yields much, but it hardly brings us nearer to the secret of the Old One. In any case, I am convinced that He does not throw dice.

Thus, Einstein refused to accept probability-based laws as final in physics, and here for the first time he spoke of the dice-playing God—an expression he used many times later as he expressed his belief that there are deterministic laws yet to be found. Despite the refusal of Einstein (and others) to accept probability laws in mechanics, neither he

Tempting as it is for us old-timers, it is probably not wise to harp on the paradox of wave and particle properties, saddling the young with our own limitations of intuition. Perhaps new generations ought to be be brought up to feel comfortable with relativity and quantum mechanics, rather than continually reminded of its peculiarity from the classical view.

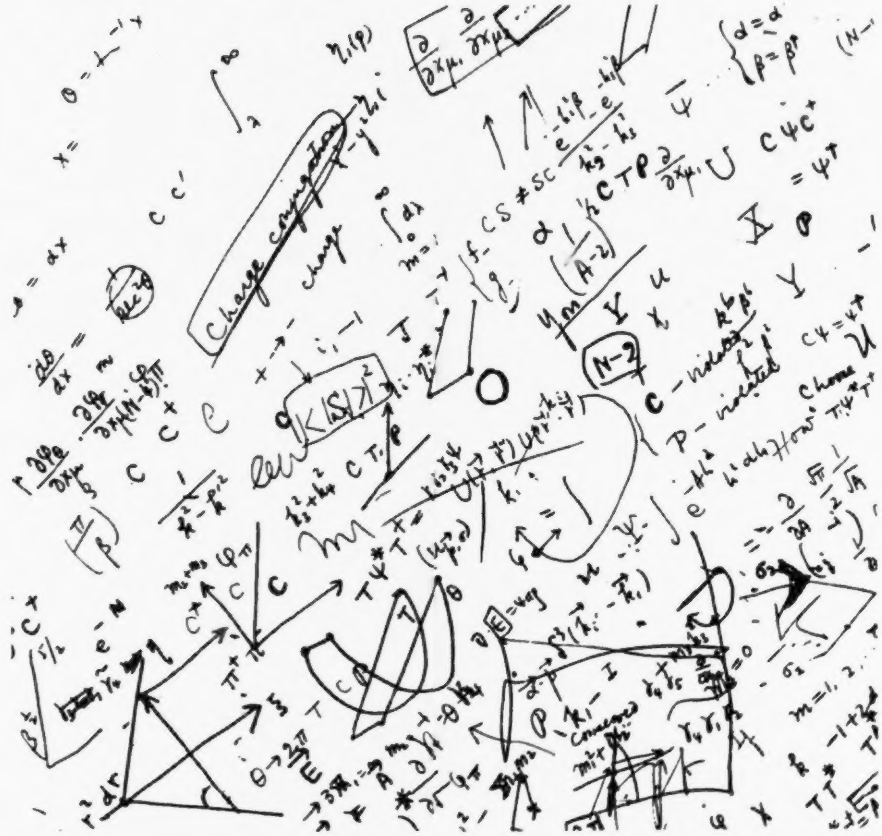
nor any other physicist has succeeded in replacing Born's probability interpretation of quantum mechanics.

Scientists agree that quantum mechanics works; it gives the right answers to many questions in physics, it unifies ideas and occurrences that were once unconnected, and it has been wonderfully productive of new experiments and new concepts. On the other hand, there is less agreement about the significance of quantum theory. Quantum theory yields probability functions, not particle trajectories. Some scientists see in this aspect of the theory an important revelation about the nature of the world; for other scientists this same fact indicates that quantum theory is incomplete. Some in this second group are trying to develop a more basic, non-statistical theory for which the present quantum theory is only a limiting case. There is no doubt that quantum theory has profoundly influenced man's views of nature. It would be a mistake to assume that quantum mechanics provides some sort of ultimate physical theory, although up to this time no one has developed a successful nonstatistical theory of atomic and nuclear physics.

Finally, it must be stressed again that effects which are completely unnoticeable because of the large masses of the visible world are very important for the small particles of the atomic world. The simple concepts (such as wave, particle, position, velocity) which work satisfactorily for the world of everyday experience are not appropriate, and the attempt to borrow these concepts for the atomic world has produced our problems of interpretation. We have been lucky enough to have unscrambled many of the apparent paradoxes, although we may at first be unhappy to have lost a world in which waves were only waves and particles were only particles.

Q11 In wave terms, the bright lines of a diffraction pattern are regions where there is a high field intensity produced by constructive interference. In the probability interpretation of quantum mechanics, the bright lines of a diffraction pattern are regions where there is a high ....?

Q12 If quantum mechanics can predict only probabilities for the behavior of any one particle, how can it predict many phenomena, for example, half-lives and diffraction patterns, with great certainty?


Charge conjugation
change
C - violated
P - violated
Choose
CTP
CS ≠ SC
C C'
C ψ C⁺
= ψ⁺
N-2
|K|S|K|²

Doodles from the scratch pad of a modern theoretical physicist, Prof. C. N. Yang.

C. N. Yang, along with T. D. Lee, was awarded the Nobel Prize in 1957.

4.2×10^7 m/sec

20.1 How fast would you have to move to increase your mass by 1%?

20.2 The centripetal force on a mass moving with relativistic speed v around a circular orbit of radius R is $F = mv^2/R$, where m is the relativistic mass. Electrons moving at a speed 0.60 c are to be deflected in a circle of radius 1.0 m: what must be the magnitude of the force applied? ($m_o = 9.1 \times 10^{-31}$ kg.) 3.7×10^{-14} N

20.3 The formulas ($p = m_o v$, $KE = \frac{1}{2}mv^2$) used in Newtonian physics are convenient approximations to the more general relativistic formulas. The factor $1/\sqrt{1-v^2/c^2}$ can be expressed as an infinite series of steadily decreasing terms by using a binomial series expansion. When this is done we find that

$$1/\sqrt{1-v^2/c^2} = 1 + 1/2\,\frac{v^2}{c^2} + 3/8\,\frac{v^4}{c^4} + 5/16\,\frac{v^6}{c^6} + 35/128\,\frac{v^8}{c^8} + \ldots.$$

a) Show, by simple substitution, that when $\frac{v}{c}$ is less than 0.1, the values of the terms drop off so rapidly that only the first few terms need be considered.

b) The greatest speeds of man-sized objects are rarely more than 3,000 m/sec, so $\frac{v}{c}$ is less than 10^{-5}. Substitute the series expression for $1/\sqrt{1-v^2/c^2}$ into the relativistic formulas,

$$p = \frac{m_o v}{\sqrt{1-v^2/c^2}} \quad \text{and}$$

$$KE = mc^2 - m_o c^2$$

and cross off terms which will be too small to be measurable. What formula would you use for momentum and kinetic energy in describing the motion of man-sized objects? $KE = \frac{1}{2}\, m_o v^2$

20.4 According to relativity theory, changing the energy of a system by ΔE also changes the mass of the system by $\Delta m = \Delta E/c^2$. Something like 10^5 joules per kilogram of substance are commonly released as heat energy in chemical reactions.

a) Why then aren't mass changes detected in chemical reactions? too small

b) Calculate the mass change associated with a change of energy of 10^5 joules. 1.1×10^{-12} kg

20.5 The speed of the earth in its orbit is about 18 miles/sec (3×10^4 m/sec). Its "rest" mass is 6.0×10^{24} kg.

a) What is the kinetic energy ($\frac{1}{2}m_o v^2$) of the earth in its orbit? 27×10^{32} J

b) What is the mass equivalent of that kinetic energy? 3.0×10^{16} kg

c) By what percentage is the earth's "rest" mass increased at orbital speed? 5×10^{-7} %

d) Refer back to Unit 2 to recall how the mass of the earth is found; was it the rest mass or the mass at orbital speed? rest mass

20.6 In 1926, Sir John Squire proposed the following continuation of Pope's verse on Newton:

Pope: Nature and Nature's laws lay hid in night
God said, 'Let Newton be!' and all was light.

Squire: It did not last: The Devil howling 'Ho,
Let Einstein be,' restored the status quo.

What does this mean, and do you agree? discussion

20.7 In relativistic mechanics the formula $\vec{p} = m\vec{v}$ still holds, but the mass m is given by $m = m_o/\sqrt{1-v^2/c^2}$. The rest mass of an electron is 9.1×10^{-31} kg.

a) What is its momentum when it is moving down the axis of a linear accelerator from left to right at a speed of 0.4 c with respect to the accelerator tube? 1.2×10^{-22} kg·m/sec

b) What would Newton have calculated for the momentum of the electron? 1.1×10^{-22} kg·m/sec

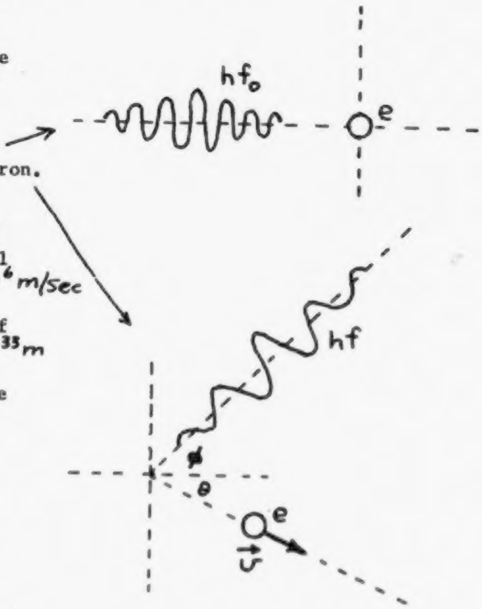
c) By how much would the relativistic momentum increase if the speed of the electron were doubled? 2.4×10^{-22} kg·m/sec

d) What would Newton have calculated its change in momentum to be? 1.1×10^{-22} kg·m/sec

20.8 Calculate the momentum of a photon of wavelength 4000 Å. 1.7×10^{-27} kg·m/sec How fast would an electron have to move in order to have the same momentum? 1.9×10^{3} m/sec

20.9 What explanation would you offer for the fact that the wave aspect of light was shown to be valid before the particle aspect was demonstrated? discussion

20.10 Construct a diagram showing the change that occurs in the frequency of a photon as a result of its collision with an electron.



20.11 The electrons which produced the diffraction photograph on p. 109 had de Broglie wavelengths of 10^{-10} meter. To what speed must they have been accelerated? (Assume that the speed is small compared to c, so that the electron mass is $\approx 10^{-30}$ kg.) 6.6×10^{6} m/sec

20.12 A billiard ball of mass 0.2 kilograms moves with a speed of 1 meter per second. What is its de Broglie wavelength? 3.3×10^{-33} m

20.13 Show that the de Broglie wavelength of a classical particle of mass m and kinetic energy KE is given by

$$\lambda = \frac{h}{\sqrt{2m(KE)}}$$

What happens when the mass is very small and the speed is very great? λ becomes larger

20.14 Suppose that the only way you could obtain information about the world was by throwing rubber balls at the objects around you and measuring their speeds and directions of rebound. What kinds of objects would you be unable to learn about? discussion

20.15 A bullet can be considered as a particle having dimensions approximately 1 centimeter. It has a mass of about 10 grams and a speed of about 3×10^{4} centimeters per second. Suppose we can measure its speed to within one part of 10^{4}. What is the corresponding uncertainty in its position according to Heisenberg's principle? 3×10^{-31} m

20.16 Show that if Planck's constant were equal to zero, quantum effects would disappear and particles would behave according to Newtonian physics. What effect would this have on the properties of light? discussion

20.17 Bohr once said,

> If one does not feel a little dizzy when discussing the implications of Planck's constant h it means that one does not understand what one is talking about.

What might he have meant? (Refer to examples from Chapters 18, 19 and 20.) Do you agree with Bohr's reaction? discussion

20.18 A particle confined in a box cannot have a kinetic energy less than a certain amount; this least amount corresponds to the longest de Broglie wavelength which produces standing waves in the box; that is, the box size is one-half wavelength. For each of the following situations find the longest de Broglie wavelength that would fit in the box; then use $p = h/\lambda$ to find the momentum p, and use $p = mv$ to find the speed v.

a) a dust particle (about 10^{-9} kg) in a display case (about 1 m across).

b) an argon atom (6.6×10^{-26} kg) in a light bulb (about 10^{-1} m across).

c) a protein molecule (about 10^{-22} kg) in a bacterium (about 10^{-6} m across).

d) an electron (about 10^{-30} kg) in an atom (about 10^{-10} m across).

20.18

a $\lambda = 2\ m$
$p = 3.3 \times 10^{-34}\ kg \cdot m/sec$
$v = 3.3 \times 10^{-25}\ m/sec$

b $v = 5 \times 10^{-8}\ m/sec$

c $v = 3.3 \times 10^{-6}\ m/sec$

d $v = 3.3 \times 10^{6}\ m/sec$

20.19 Some philosophers (and some physicists) have claimed that the Uncertainty Principle proves that there is free will. Do you think this extrapolation from atomic phenomena to the world of animate beings is justified? Discuss.

discussion

20.20 A physicist has written

> It is enough that quantum mechanics predicts the average value of observable quantities correctly. It is not really essential that the mathematical symbols and processes correspond to some intelligible physical picture of the atomic world.

discussion

Do you regard such a statement as acceptable? Give reasons.

20.21 The great French physicist Pierre Laplace (1748-1827) wrote,

> Given for one instant an intelligence which could comprehend all the forces by which nature is animated and the respective situation of the beings who compose it—an intelligence sufficiently vast to submit these data to analysis—it would embrace in the same formula the movements of the greatest bodies of the universe and those of the lightest atom; for it, nothing would be uncertain and the future, as the past, would be present to its eyes. A Philosophical Essay on Probabilities.

discussion

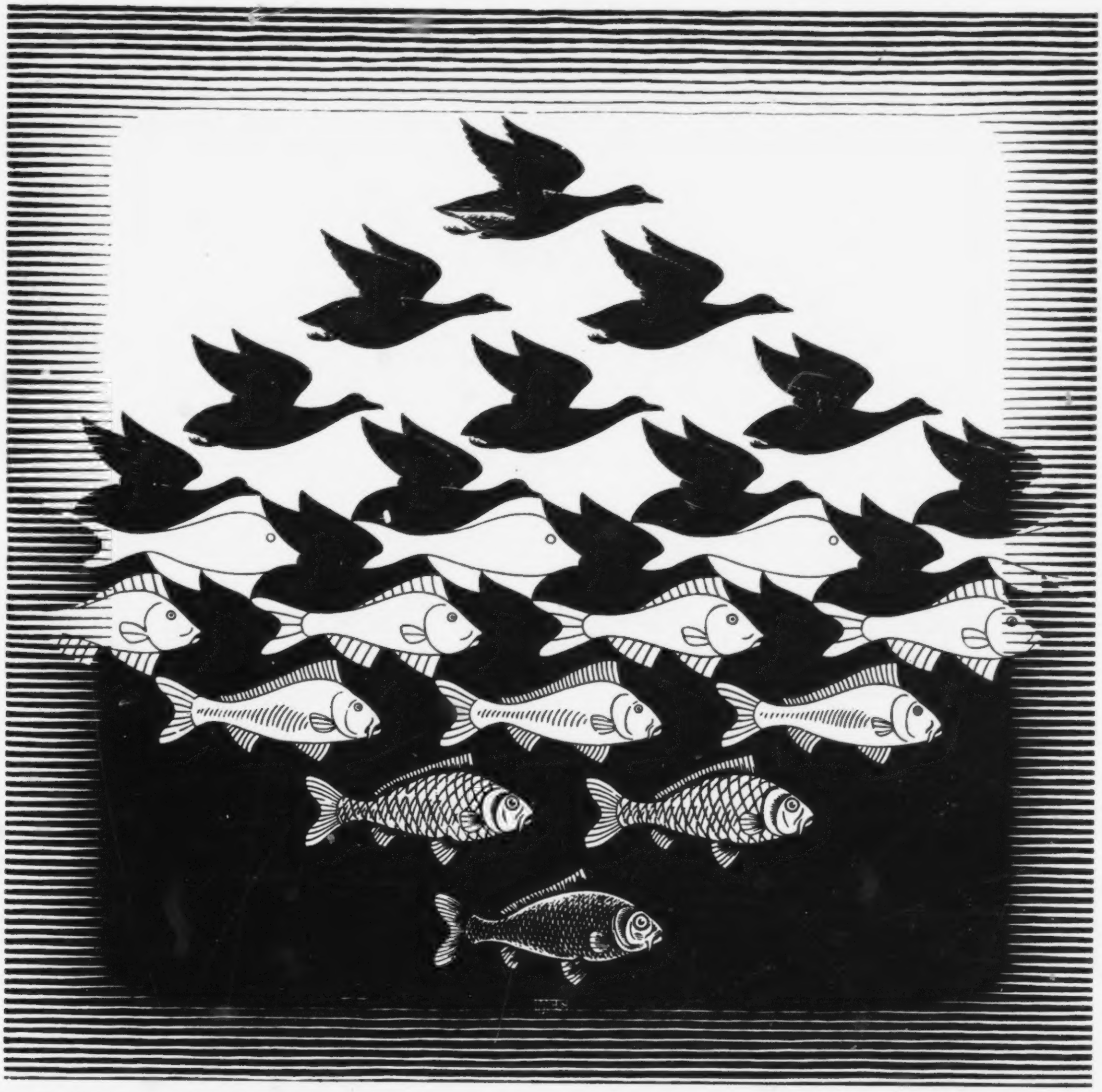
Is this statement consistent with modern physical theory?

20.22 In Chapters 19 and 20 we have seen that it is impossible to avoid the wave-particle dualism of light and matter. Bohr has coined the word complementarity for the situation in which two opposite views seem equally valid, depending on which aspect of a phenomenon one chooses to consider. Can you think of situations in other fields (outside of atomic physics) to which this idea might apply?

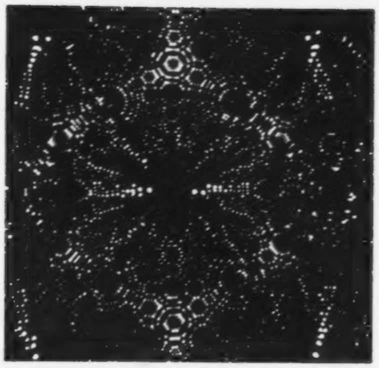
discussion

20.23 In Units 1 through 4 we discussed the behavior of large-scale "classical particles" (for example, tennis balls) and "classical waves" (for example, sound waves), that is, of particles and waves that in most cases can be described without any use of ideas such as the quantum of energy or the de Broglie matter-wave. Does this mean that there is one sort of physics ("classical physics") for the phenomena of the large-scale world and quite a different physics ("quantum physics") for the phenomena of the atomic world? Or does it mean that quantum physics really applies to all phenomena but is not distinguishable for classical physics when applied to large-scale particles and waves? What arguments or examples would you use to defend your answer?

discussion



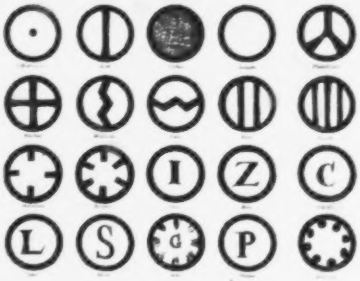
M. C. Escher, "Sea and Sky."



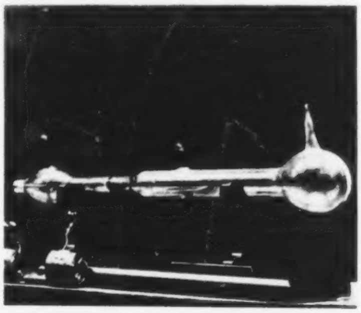
Epilogue We have traced the concept of the atom from the early ideas of the Greeks to the quantum mechanics now generally accepted by physicists. The search for the atom started with the qualitative assumptions of Leucippus and Democritus who thought that their atoms offered a rational explanation of things and their changes. For many centuries most natural philosophers thought that other explanations, not involving atoms, were more reasonable. Atomism was pushed aside and received only occasional consideration until the seventeenth century. With the growth of the mechanical philosophy of nature in the seventeenth and eighteenth centuries, particles (corpuscles) became important. Atomism was reexamined, mostly in connection with physical properties of matter. Boyle, Newton and others speculated on the role of particles in the expansion and contraction of gases. Chemists speculated about atoms in connection with chemical change. Finally, Dalton began the modern development of atomic theory, introducing a quantitative aspect that had been lacking—the relative atomic mass.

Chemists, in the nineteenth century, found that they could correlate the results of many chemical experiments in terms of atoms and molecules. They also found a system in the properties of the chemical elements. Quantitative information about atomic masses provided a framework for the system—the periodic table of Mendeleev. During the nineteenth century, physicists developed the kinetic theory of gases. This theory—based on the assumption of very small corpuscles, or particles, or molecules, or whatever else they might be called—helped strengthen the position of the atomists. Other work of nineteenth-century physics helped pave the way to the study of the structure of atoms, although the reasons for this work had no direct connection with the problem of atomic structure. The study of the spectra of the elements and of the conduction of electricity in gases, the discovery of cathode rays, electrons, and x rays, all eventually led to the atom.

Nineteenth-century chemistry and physics converged, at the beginning of the twentieth century, on the problem of atomic structure. It became clear that the uncuttable, infinitely hard atom was too simple a model: that the atom itself is made up of smaller particles. And so the search for a model with structure began. Of the early models, that of Thomson—the pudding with raisins in it—attracted much interest; but it was inadequate. Then came Rutherford's nuclear atom, with its small, heavy, positively charged nucleus, surrounded, somehow, by negative charges. Then the atom of Bohr, with its electrons moving in orbits like planets in a miniature solar system. The Bohr theory had many successes and linked chemistry and spectra to the physics of atomic structure. But then the Bohr theory fell, and with it the easily grasped pictures of the atom. There is an end—at least for the present—to the making of simple physical models. Now is the time for mathematical models, for equations, not for pictures. Quantum mechanics enables us to calculate how atoms behave: it helps us understand the physical and chemical properties of the elements. What we used to call "atomic physics," Dirac now calls "the theory of chemistry," presumably because "chemistry" is that which is understood, while physics still has secrets.

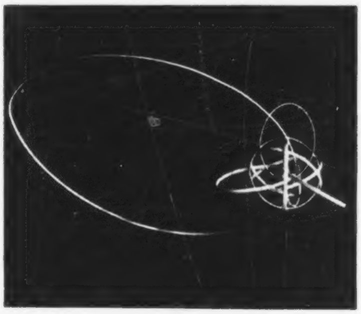


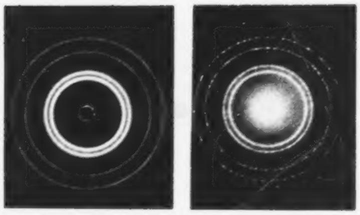
The next stage in our story is the nucleus of the atom. Is it uncuttable? Is it infinitely hard? Or is the nucleus made up of smaller components? Do we have to worry about its composition and structure?



Is the problem of matter like that
of the so-famous fleas?

Big fleas have little fleas
Upon their backs to bite 'em,
And little fleas have lesser fleas,
And so ad infinitum.





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Brief Answers to Study Guide

Chapter 17

17.1	80.3% zinc 19.7% oxygen
17.2	47.9% zinc
17.3	13.9 × mass of H atom same
17.4	986 g nitrogen 214 g hydrogen
17.5	9.23 × mass of H atom
17.6	(a) 14.1 (b) 28.2 (c) 7.0
17.7	Discussion
17.8	Na:1 Ca:2 Al:3 Sn:4 P:5
17.9	(a) Ar-K Co-Ni Te-I Th-Pa U-Np Es-Fm Md-No (b) Discussion
17.10	Discussion
17.11	Discussion
17.12	0.113 g hydrogen 0.895 g oxygen
17.13	(a) 0.05 g zinc (b) 0.30 g zinc (c) 1.2 g zinc
17.14	(a) 0.88 g chlorine (b) 3.14 g iodine (c) Discussion
17.15	Discussion
17.16	Discussion
17.17	Discussion
17.18	Discussion

Chapter 18

18.1	(a) 2.0×10^7 m/sec (b) 1.8×10^{11} coul/kg
18.2	Proof
18.3	Discussion
18.4	Discussion
18.5	1.5×10^{14} cycles/sec ultraviolet
18.6	4×10^{-19} joules 4×10^{-18} joules
18.7	2.6×10^{-19} joules 1.6 eV
18.8	4.9×10^{14} cycles/sec
18.9	(a) 2.5×10^{20} photons (b) 2.5 photons/sec (c) 0.4 sec (d) 2.5×10^{-10} photons (e) 0.1 amp
18.10	1.3×10^{17} photons
18.11	1.2×10^{19} cycles/sec
18.12	Discussion
18.13	1.2×10^5 volts 1.9×10^{-14} joules 1.2×10^5 eV
18.14	Glossary
18.15	Discussion

Chapter 19

19.1	Discussion
19.2	Five listed in text, but theoretically an infinite number. Four lines in visible region.
19.3	$n_i = 8$ 3880 Å $n_i = 10$ 3790 Å $n_i = 12$ 3730 Å Discussion
19.4	(a) Discussion (b) $n_i = \infty$ (c) Lyman series 910 Å Balmer series 3650 Å Paschen series 8200 Å (d) $E = 21.8 \times 10^{-19}$ joules $E = 13.6$ eV
19.5	Discussion
19.6	Discussion
19.7	Discussion
19.8	Ratio = 10^{-4}
19.9	3.5 meters
19.10	Discussion
19.11	Discussion
19.12	Discussion
19.13	Discussion
19.14	Discussion
19.15	$Z = 36$, $Z = 54$
19.16	Glossary
19.17	Discussion
19.18	Discussion

Chapter 20

20.1 0.14 c or 4.2×10^{7} m/sec

20.2 3.7×10^{-14} newtons

20.3 $m_o v^2$ and $m_o v$

20.4 (a) changes are too small
(b) 10^{-12} kg

20.5 (a) 27×10^{32} joules
(b) 3×10^{16} kg
(c) 5×10^{-7} %
(d) Rest mass

20.6 Discussion

20.7 (a) 1.2×10^{-22} kg·m/sec
(b) 1.1×10^{-22} kg·m/sec
(c) 2.4×10^{-22} kg·m/sec
(d) 1.1×10^{-22} kg·m/sec

20.8 $p = 1.7 \times 10^{-27}$ kg·m/sec
$v = 1.9 \times 10^{3}$ m/sec

20.9 Discussion

20.10 Diagram

20.11 6.6×10^{6} m/sec

20.12 3.3×10^{-33} m

20.13 Proof

20.14 Discussion

20.15 $\Delta x = 3.3 \times 10^{-31}$ m

20.16 Discussion

20.17 Discussion

20.18 (a) 2 m, $3.3 \times 10^{-24}\ \frac{kg\cdot m}{sec}$, $3.3 \times 10^{-25}\ \frac{m}{sec}$
(b) 0.2 m, $3.3 \times 10^{-33}\ \frac{kg\cdot m}{sec}$, $5 \times 10^{-8}\ \frac{m}{sec}$
(c) 2×10^{-6} m, $3.3 \times 10^{-28}\ \frac{kg\cdot m}{sec}$, $3.3 \times 10^{-6}\ \frac{m}{sec}$
(d) 2×10^{-10} m, $3.3 \times 10^{-24}\ \frac{kg\cdot m}{sec}$, $3.3 \times 10^{6}\ \frac{m}{sec}$

20.19 Discussion

20.20 Discussion

20.21 Discussion

20.22 Discussion

20.23 Discussion

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Prologue

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P. 9 Diderot, Denis, Encyclopédie. Houghton Library, Harvard University.

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Answers to End of Section Questions

Chapter 17

Q1 The atoms of any one element are identical and unchanging.

Q2 conservation of matter; the constant ratio of combining weights of elements

Q3 no

Q4 It was the highest known element—and others were rough multiples.

Q5 relative mass; and combining number, or "valence"

Q6 2,3,6,1,2

Q7 density, melting point, chemical activity, valence

Q8 atomic mass

Q9 when the chemical properties clearly suggested a slight change or order

Q10 Sometimes the next heavier element didn't have the expected properties—but did have the properties for the next space over.

Q11 its position in the periodic table, determined by many properties but usually increasing regularly with atomic mass

Q12 Water, which had always been considered a basic element, and had resisted all efforts at decomposition, was easily decomposed.

Q13 New metals were separated from substances which had never been decomposed before.

Q14 the amount of charge transferred by the current, the valence of the elements, and the atomic mass of the element

Chapter 18

Q1 They could be deflected by magnetic and electric fields.

Q2 because the mass is 1800 times smaller

Q3 (1) Identical electrons were emitted by a variety of materials; and (2) the mass of an electron was much smaller than that of an atom.

Q4 All other values of charge he found were multiples of that lowest value.

Q5 Fewer electrons are omitted, but with the same average energy as before.

Q6 The average kinetic energy of the emitted electrons decreases until, below some frequency value, none are emitted at all.

Q7 The energy of the quantum is proportional to the frequency of the wave, $E = hf$.

Q8 The electron loses some kinetic energy in escaping from the surface.

Q9 The maximum kinetic energy of emitted electrons is 2.0 eV.

Q10 When x rays passed through material, say air, they caused electrons to be ejected from molecules, and so produced + ions.

Q11 (1) not deflected by magnetic field; (2) show diffraction patterns when passing through crystals; (3) produced a pronounced photoelectric effect

Q12 (1) diffraction pattern formed by "slits" with atomic spacing (that is, crystals); (2) energy of quantum in photoelectric effect

Q13 For atoms to be electrically neutral, they must contain enough positive charge to balance the negative charge of the electrons they contain; but electrons are thousands of times lighter than atoms.

Chapter 19

Q1 It is composed of only certain frequencies of light.

Q2 by heating or electrically exciting a gas (However, very dense gas, such as the insides of a star, will emit a continuous range of light frequencies.)

Q3 Certain frequencies of light are missing.

Q4 by passing light with complete range of frequencies through a relatively cool gas

Q5 none (he predicted that they would exist because the mathematics was so neat.)

Q6 careful measurement and tabulation of data on spectral lines

Q7 They have a positive electric charge and are repelled by the positive electric charge in atoms.

Q8 Rutherford's model located the positively charged bulk of the atom in a tiny nucleus—in Thomson's model the positive bulk filled the entire atom.

Q9 the number of positive electron charges in the nucleus

Q10 3 positive units of charge (when all 3 electrons were removed)

Q11 Atoms of a gas emit light of only certain frequencies, which implies that each atom's energy can change only by certain amounts.

Q12 none (He assumed that electron orbits could have only certain value, of angular momentum, which implied only certain energy states.)

Q13 Bohr derived his prediction from a physical model, from which other predictions could be made. Balmer only followed out a mathematical analogy.

Q14 (a) 4.0 eV (b) 0.1 eV (c) 2.1 eV

Q15 The electron arrangements in noble gases are very stable. When an additional nuclear charge and an additional electron are added, the added electron is bound very weakly to the atom.

Q16 It predicted some results that disagreed with experiment; and it predicted others which could not be tested in any known way.

Chapter 20

Q1 It increases, without limit.

Q2 It increases, approaching ever nearer to a limiting value, the speed of light.

Q3 Photon momentum is directly proportional to the frequency of the associated wave.

Q4 That the idea of photon momentum is consistent with the experimental results of scattering of x rays by electrons.

Q5 by analogy with the same relation for photons

Q6 The regular spacing of atoms in crystals is about the same as the wavelength of low-energy electrons.

Q7 Bohr invented his postulate just for the purpose. Schrödinger's equation was derived from the wave nature of electrons and explained many phenomena other than hydrogen spectra.

Q8 It is almost entirely mathematical—no physical picture or models can be made of it.

Q9 It can. But less energetic photons have longer associated wavelengths, so that the location of the particle becomes less precise.

Q10 It can. But the more energetic photons will disturb the particle more and make measurement of velocity less precise.

Q11 ...probability of quanta arriving.

Q12 As with all probability laws, the average behavior of a large collection of particles can be predicted with great precision.

Acknowledgments

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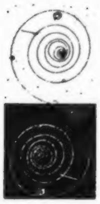
Project Physics **Teacher's Guide**

An Introduction to Physics 5 Models of the Atom



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Overview of Unit 5

Evidence which supports an atomic theory of matter is presented in the first half of this unit. There is no single experiment upon which atomic theory is based. Rather, a number of experimental data, like the interlocking pieces of a puzzle, provide the basis for the theory.

The internal structure of the atom is the subject of Chapters 18 and 19. The experiments of Thomson, Millikan and Rutherford, along with Einstein's quantum interpretation of the photoelectric effect, set the stage for the Bohr model of the atom. The Bohr model was successful in correlating much of the data that had accumulated by 1913.

Chapter 20 surveys the quantum theory of matter which followed the failure of the Bohr theory.

Experiments

E39	The charge-to-mass ratio for an electron
E40	The measurement of elementary charge
E41*	Electrolysis
E42*	Photoelectric effect
E43*	Spectroscopy

Transparencies

T35	Periodic table
T36	Photoelectric mechanism
T37	Photoelectric equation
T38	Alpha scattering
T39	Energy levels—Bohr theory

Demonstrations

D53	Electrolysis of water
D54	Charge-to-mass ratio for cathode rays
D55	Photoelectric effect
D56	Blackbody radiation
D57	Absorption
D58	Ionization Potential

Films

F35	Definite and multiple proportions
F36	Elements, compounds and mixtures
F37	Counting electrical charges in motion
F38	Millikan experiment
F39	Photoelectric effect
F40	The structure of atoms
F41	Rutherford atom
F42	A new reality
F43	Franck-Hertz experiment
F44	Interference of photons
F45	Matter waves
F46	Light: wave and quantum theories

Loops

L46	Production of sodium by electrolysis
L47	Thomson model of the atom
L48	Rutherford scattering

Reader Articles

R1	Failure and Success
R2	Structure, Substructure, Superstructure
R3	The Island of Research
R4	The 'Thomson' Atom
R5	Einstein
R6	Mr. Tompkins and Simultaneity
R7	Mathematics and Relativity
R8	Relativity
R9	Parable of the Surveyors
R10	Outside and Inside the Elevator
R11	Einstein and Some Civilized Discontents
R12	The Teacher and the Bohr Theory of the Atom
R13	The New Landscape of Science
R14	The Evolution of the Physicist's Picture of Nature
R15	Dirac and Born
R16	I Am This Whole World: Erwin Schrödinger
R17	The Fundamental Idea of Wave Mechanics
R18	The Sentinel
R19	The Sea-Captain's Box
R20	Space Travel: Problems of Physics and Engineering
R21	Looking for a New Law

Multi-Media

1 LAB STATIONS Law of Multiple Proportions	2 PSSC FILM # 0110 (30 min) Law of Definite and Multiple Proportions.	3 DEMONSTRATION LAB E41 Faraday's Law	4 CHEM STUDY FILM Chemical Families S.G.D. on Film
1a Text: Prologue, 17.1, 17.2	2a Text: 17.3, 17.4	3a Text: 17.5, 17.6	4a Text: 17.7, 17.8
5 LECTURE-DEMONSTRATION Cathode-Ray Tubes	6 LAB E40 The Measurement of Elementary Charge	7 PSSC FILM # 0404 Millikan Experiment	8 LAB STATIONS Photoelectric Effect
5a Text: 18.1 → 18.3	6a Reader: The "Thomson" Atom, Thomson	7a Text: 18.4, 18.5	8a Text 18.6
9 DISCUSSION Photoelectric Effect	10 TEACHER PRESENTATION X rays	11 LAB STATIONS Observing Spectra	12 DEMONSTRATION EXPERIMENT Spectra
9a Text: 19.1 → 19.3	10a Reader: A Remarkable Regularity..., Balmer	11a Text: 19.4 → 19.7	12a Text: 19.8, 19.9
13 TEACHER PRESENTATION Models of the Atom	14 PSSC FILM # 0416 (40 min) Rutherford Atom	15 TEACHER PRESENTATION Bohr Model of Hydrogen Atom	16 TEACHER PRESENTATION Where Bohr Fails...
13a Text: 20.1, 20.2	14a Text: 20.3, 20.4	15a Reader: The Teacher and..., C.P. Snow	16a Text: 20.5, 20.6
17 SMALL GROUP RESEARCH	18 SMALL GROUP RESEARCH Continued	19 Groups prepare presentations	20 CLASS PRESENTATIONS of special topic
17a Special Reading	18a →	19a →	20a →
21 PRESENTATIONS Continued	22 REVIEW UNIT	23 UNIT TEST	24
21a	22a	23a	24a

Details of the Multi-Media Schedule

Day 1

Lab stations: Law of Multiple Proportions

1. mechanical models of chemical compounds (See CHEM study materials.)
2. other mechanical models
3. film loop #46, Production of Sodium by Electrolysis
4. Pass current through very weak H_2SO_4 solution. Measure ratio of volumes of gases produced.
5. Dalton's Puzzle (See Student Handbook.)
6. Cigar-box "molecules" (See Student Handbook or CHEM Study Experiment.)

Day 2

Film: Definite and Multiple Proportions (PSSC) #0110 (30 min)

This film is used to tie together ideas introduced on Day 1 and hence does not require elaborate follow-up discussion.

The rest of the class period is spent solving problems. See end of Chapter 17 for examples.

Day 3

Experiment 41: Electrolysis of Metals

Teacher or students gather data in advance. Experiment is shown qualitatively during class and calculations made from previous data.

Day 5

Lecture demonstration of cathode rays and Thomson q_e/m experiment. Treatment will vary with equipment available. A good ending for the class might be the film loop on the Thomson atom (HPP #47).

Day 6

Experiment 40: Measurement of Elementary Charge

Have Millikan apparatus set up in advance. Students spend first 15 to 20 minutes with apparatus. Since adequate data is difficult to obtain in a short time, data from previous experiments or from Teacher Guide may be given out. Students try to find q.

Day 7

Film: Millikan Experiment PSSC #0404

Show first 15 minutes of the film up to point where charge is changed by x-ray bombardment. Stop the film at that point and discuss briefly.

Teacher presentation: Introduction to Photoelectric Effect

Day 8

Lab Stations: Photoelectric Effect (Experiment 42)

Set up 3 or 4 stations using Project Physics amplifiers and photoelectric tubes. Half the class spends half the period gathering data.

Set up 3 or 4 stations using an electroscope and zinc plate. Charge the electroscope and shine light on the plate. Measure the rate of discharge. (A similar experiment is written up in PSSC.) Half the class spends half the hour on this, then rotates to Project Physics apparatus.

Day 11

Lab Stations: Observing Spectra

Students observe and make notes describing spectra from
1) incandescent lamp
2) neon lamp
3) hydrogen capillary tube
4) helium capillary tube
5) sodium vapor in flame
6) fluorescent tube

Day 12

Spectra continued

Students are issued photographs (See Experiment 42) of spectra and asked to calculate wavelengths and identify elements. Photographs may be taken in advance by an interested student.

Day 13

Teacher presentation
1) Thomson model of the atom
2) Rutherford's experiment
3) Rutherford's model of the atom

See Teacher Guide for demonstrations.

Day 17

Small-group library research on topics related to Unit 5

Some examples:

1) special theory of relativity
2) De Broglie waves
3) wave-particle dualism
4) Compton effect
5) uses of spectra in astronomy

Day 19

Small groups prepare presentations of research material to class.

Encourage students to effectively communicate ideas that they have researched. Dramatizations, readings, and use of audio visual aids will add interest, and allow all students to participate.

Chapter 17 Schedule Blocks

Each block represents one day of classroom activity and implies a 50-minute period. The words in each block indicate only the basic material under consideration.

Chapter 17: The Chemical Basis of Atomic Theory

Read Prologue, 17.1, 17.2

Early atomic theories

Read 17.3 → 17.5

Chemical properties and the periodic table

Read 17.6 – 17.7

Synthesis of electricity and matter

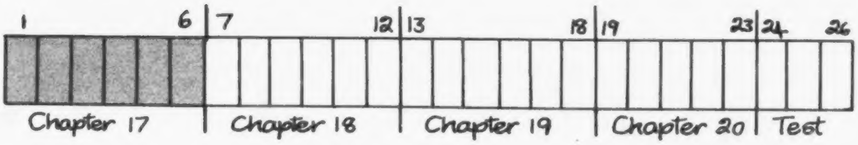
Read 17.8 and E41*

Lab
Electrolysis effects

Post Lab and/or problem seminar

Review

Test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26
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Chapter 18
Chapter 19
Chapter 20
Test

Chapter 17 Resources Chart

Text	[illegible]	Problems			Experiments and Further Demonstrations
Prologue					
17.1 Dalton's atomic theory and the laws of chemical combination		1 2	3		
17.2 The atomic masses of the elements		5	4 6 7		
17.3 Other properties of the elements: valence		8			
17.4 The search for order and regularity among the elements:					
17.5 Mendeleev's periodic table of the elements					
17.6 The modern periodic table			9	10 11	
17.7 Electricity and matter: qualitative studies					D53 Electrolysis of water E 41 * Electrolysis effect
17.8 Electricity and matter: quantitative studies		15 16 17	12 13 14 18		

Chapter 17 Resources Chart

Transparencies, Loops, Films, Reader, Programmed Instruction	Student Activities
Structure, Substructure, Superstructure	
F35 Definite and mutiple proportions F36 Elements, compounds, and mixtures R1 Failure and Success	
	Dalton's Puzzle
R21 Looking For a New Law	
T35 Periodic table	Periodic Table
F37 Counting electrical charges in motion	The electrolysis of water Single-electrode plating
L46 Producton of sodium by electrolysis	

Summaries and Equipment: Unit 5

Some of the equipment used in these experiments is more expensive than that used in earlier units. The Project is not able to supply enough equipment to enable each group of students to do the experiment at the same time. It will therefore be necessary to "rotate" the students, so that some are using the Millikan apparatus while others are doing the experiment on the photoelectric effect, and others are photographing the Balmer spectrum.

E41*: Electrolysis Effect

The cathode is suspended from the arm of a balance, and its increase in mass is determined without removing it from the plating bath. Students already know the value of q_e from the previous experiment. The results of this experiment—mass of copper deposited when a known quantity of charge passes—can be used to calculate the mass of a single copper atom.

Equipment (for details see Teacher Guide notes to the experiment itself)

500 or 600 ml beaker

Copper sheet

Balance—equal arm or triple beam

Saturated solution of copper sulfate in distilled water, with two or three drops of concentrated sulfuric acid

6-volt 5-amp dc power supply

Ammeter 0-5 or 0-10 amps dc

Rheostat or variable autotransformer

Fine copper wire and clips

Hook-up wire

Stopwatch (optional)

Chapter 18 Schedule Blocks

Each block represents one day of classroom activity and implies a 50-minute period. The words in each block indicate only the basic material under consideration.

Chapter 18: Electrons and Quanta

Read 18.1, 18.2

Discovery
of the
Electron

Read 18.3 and E40*

Lab –
Measurement
of elementray
charge

Post lab
and/or
problem seminar

Read 18.4, 18.5

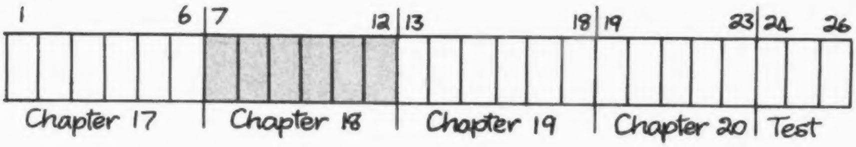
Photoelectric
effect

Read 18.6, 18.7

X Rays
atomic models

Review

Test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Chapter 17
Chapter 18
Chapter 19
Chapter 20
Test

Chapter 18 Resources Chart

Text	Study Guide	Problems			Experiments and Teacher Demonstrations
18.1 The problem of atomic structure: pieces of atoms					
18.2 Cathode rays			1		D54 Charge-to-mass ratio for cathode rays E39 The charge-to-mass ratio for an electron
18.3 The measurement of the charge of the electron: Millikan's experiment		2			E40 * The measurement of elementary charge
18.4 The photoelectric effect		3			D55 Photoelectric effect E42 * Photoelectric effect
18.5 Einstein's theory of the photoelectric effect: quanta		4 5 6 8	7 10	9 15	
18.6 X rays		11	12 13		
18.7 Electrons, quanta, and the atom		14			

Chapter 18 Resources Chart

Transparencies, Loops, Films, Reader, Programmed Instruction	Student Activities
'18 The Sentinel	
	Cathode rays in a Crooke's tube Measuring q/m for the electron
F38 Millikan experiment	
T36 Photoelectric mechanism	Lighting a light bulb with a match photoelectrically
T37 Photoelectric equation F39 Photoelectric effect R5 Einstein R20 Space Travel: Problems of Physics and Engineering	Writings by and about Einstein
	X-rays from a Crooke's tube
L47 Thomson model of the atom R4 The 'Thomson Atom'	

E39: The Charge-to-Mass Ratio for an Electron

Students use their "home-made" electron-beam tubes (Experiment 37) to repeat J. J. Thomson's classic experiment. Results for q_e/m should be within an order of magnitude of the accepted value.

Equipment

- Electron-beam tube that gives at least 5 cm visible beam (from E37)
- Vacuum pump, power supply, hook-up wire as in E37
- Cardboard tube, about 3" diameter and 6" long
- Copper magnet wire (anything between 18 and 28 gauge can be used)
- Current balance and all accessories (power supply, ammeter, etc.) for calibration of magnet

E40*: The Measurement of Elementary Charge (Millikan Experiment)

This modification of Millikan's experiment is used to measure the charges on tiny latex spheres electrically suspended between parallel charged plates. The charge on the electron q_e may be deduced from the data.

Equipment

- Project Physics Millikan Apparatus, complete
- Suspension of latex spheres in water
- Power supply 6V, 5A for light source
- Power supply 200 to 250V dc
- Voltmeter 0-250V dc

E42*: Photoelectric Effect

This experiment gives evidence that the wave model of light is inadequate.

The experiment shows that the kinetic energy of photoelectrons knocked out of a photosensitive surface depends on the color of the incident light.

The experiment then goes on to show that the maximum kinetic energy of the electrons is a linear function of the frequency of the incident light. A graph of energy against frequency is a straight line whose slope (as found by precise measurements) is Planck's constant h. Measurements of h in this experiment are within an order of magnitude of the accepted value h—6.62 × 10^{-34} joule·sec.

Equipment (for more details see Teacher Guide notes to the experiment itself)

- Phototube unit
- Amplifier/power-supply
- Loudspeaker, earphones or CRO (or microammeter)
- Colored filters
- Light source—mercury lamp, or fluorescent or incandescent lamp
- (voltmeter 0-2.5V dc)

Chapter 19 Schedule Blocks

Each block represents one day of classroom activity and implies a 50-minute period. The words in each block indicate only the basic material under consideration.

Chapter 19: The Rutherford-Bohr Model of the Atom

Read 19.1 and E43*

Lab
Spectroscopy

Read 19.2

Post lab discussion
Balmer relation

Read 19.3, 19.4

Rutherford's model

Read 19.5, 19.6

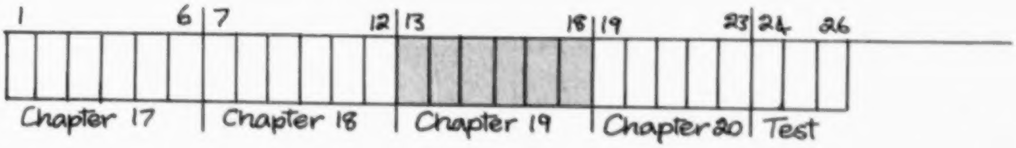
Bohr theory

Read 19.7 → 19.9

Shortcomings and prelude to a new theory

Review

Test



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Chapter 19 Resources Chart

Text	Schedule	Problems			Experiments and Teaching Demonstrations
19.1 Spectra of gases			1		D56 Blackbody radiation E43 * Spectroscopy
19.2 Regularities in the hydrogen spectrum		2	3	4	
19.3 Rutherford's nuclear model of the atom		5 6	7		
19.4 Nuclear charge and size		8			
19.5 The Bohr theory: the postulates			9		
19.6 The Bohr theory: the spectral series of hydrogen		11			
19.7 Stationary states of atoms: the Franck-Hertz experiment		12			D57 Absorption
19.8 The periodic table of the elements		14 15	13		
19.9 The failure of the Bohr theory and the state of atomic theory in the early 1920's			10 16 17 18		

Chapter 19 Resources Chart

Transparencies, Loops, Films, Reader, Programmed Instruction	Student Activities
R12 The Teacher and the Bohr Theory of the Atom	Modeling atoms with magnets Another simulation of the Rutherford atom
T38 Alpha scattering L48 Rutherford scattering F40 The structure of atoms F41 Rutherford atom	
T39 Energy levels - Bohr theory F42 A new reality	
F43 Franck-Hertz experiment	Measuring a quantum effect: ionization
T35 Periodic table	
R19 The Sea-Captain's Box	Cigar box atoms

E43*: Spectroscopy

a) Students observe as wide a variety of spectra as possible: continuous, line, absorption, etc.

b) They photograph the Balmer spectrum of hydrogen and calculate the spectral wavelength.

c) They can use these wavelengths to calculate the Rydberg constant for hydrogen, or to make an energy-level diagram for the hydrogen atom.

Equipment

Replica gratings or "spectramatics" or other "take-home" pocket spectroscopes.

Sources of line and continuous spectra such as:
a) incandescent lamp, ideally (for Part 2) supplied through a variable transformer
b) spectral tubes
c) flames, including Bunsen burner with various metallic salts added
d) fluorescent lamp
e) Balmer tube (atomic hydrogen) with power supply. If you use the Macalaster MSC1300 here be sure to remove the 6.8 megohm resistor taped to its output. Macalaster #1350 spectrum tube power supply needs no alteration.

Polaroid camera (type 002, 95, 150, or 800) and film (black-and-white, 3000 speed)

Chapter 20 Schedule Blocks

Each block represents one day of classroom activity and implies a 50-minute period. The words in each block indicate only the basic material under consideration.

Chapter 20: Some Ideas from Modern Physical Theories

Read 20.1

Mass-energy equivalence

Read 20.2

Particle behavior of waves

Read 20.3, 20.4

Wave behavior of particles

Read 20.5, 20.6

Quantum mechanics

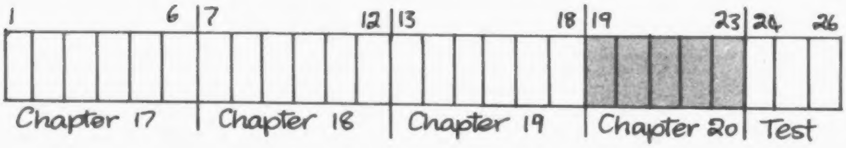
Review

Chapter test

Unit review

Unit test

Go over Unit test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Chapter 20 Resources Chart

Text	Schedule	Problems			Experiments and Teacher Demonstration
20.1 Some results of relativity theory		4 6	1 2 3 5 7		
20.2 Particle-like behavior of radiation		9 10	8		
20.3 Wave-like behavior of matter		11 12 13			
20.4 Quantum mechanics					D58 Ionization potential
20.5 Quantum mechanics – the uncertainty principle		14	15 16	17	
20.6 Quantum mechanics – probability interpretation			18 20 21 22	19	
Epilogue					

Chapter 20 Resources Chart

Transparencies, Loops, Films, Reader, Programmed Instruction	Student Activities
Mr. Tomkins and Simultaneity Mathematics and Relativity R8 Relativity R9 Parable of the Surveyors R10 Outside and Inside the Elevator	
F44 Interference of photons	
F45 Matter waves	Standing waves on a band-saw blade Turntable oscillator patterns Standing waves in a wire ring
F46 Light: wave and quantum theories R13 The New Landscape of Science R15 Dirac and Born R16 I Am This Whole World: Erwin Schrödinger	
R14 The Evolution of the Physicist's Picture of Nature	
R17 The Fundamental Idea of Wave Mechanics	
R11 Einstein and Some Civilized Discontents	

Brief Answers to Unit 5 Study Guide

17.1 80.3% zinc
19.7% oxygen

17.2 47.9% zinc

17.3 13.9 × mass of H atom
same

17.4 986 g nitrogen
214 g hydrogen

17.5 9.23 × mass of H atom

17.6 (a) 14.1
(b) 28.2
(c) 7.0

17.7 Discussion

17.8 Na:1
Ca:2
Al:3
Sn:4
P:5

17.9 (a) Ar-K
Co-Ni
Te-I
Th-Pa
U-Np
Es-Fm
Md-No
(b) Discussion

17.10 Discussion

17.11 Discussion

17.12 0.113 g hydrogen
0.895 g oxygen

17.13 (a) 0.05 g zinc
(b) 0.30 g zinc
(c) 1.2 g zinc

17.14 (a) 0.88 g chlorine
(b) 3.14 g iodine
(c) Discussion

17.15 Discussion

17.16 Discussion

17.17 Discussion

17.18 Discussion

18.1 (a) 2.0×10^{7} m/sec
(b) 1.8×10^{11} coul/kg

18.2 Proof

18.3 Discussion

18.4 Discussion

18.5 1.5×10^{14} cycles/sec
ultraviolet

18.6 4×10^{-19} joules
4×10^{-18} joules

18.7 2.6×10^{-19} joules
1.6 eV

18.8 4.9×10^{14} cycles/sec

18.9 (a) 2.5×10^{20} photons
(b) 2.5 photons/sec
(c) 0.4 sec
(d) 2.5×10^{-10} photons
(e) 0.1 amp

18.10 1.3×10^{17} photons

18.11 1.2×10^{19} cycles/sec

18.12 Discussion

18.13 1.2×10^{5} volts
1.9×10^{-14} joules
1.2×10^{5} eV

18.14 Glossary

18.15 Discussion

19.1 Discussion

19.2 Five listed in text, but theoretically an infinite number. Four lines in visible region.

19.3 $n_i = 8$ 3880 Å
$n_i = 10$ 3790 Å
$n_i = 12$ 3730 Å
Discussion

19.4 (a) Discussion
(b) $n_i = \infty$
(c) Lyman series 910 Å
Balmer series 3650 Å
Paschen series 8200 Å
(d) $E = 21.8 \times 10^{-19}$ joules
$E = 13.6$ eV

19.5 Discussion

19.6 Discussion

19.7 Discussion

19.8 Ratio = 10^{-4}

19.9 3.5 meters

19.10 Discussion

19.11 Discussion

19.12 Discussion

19.13 Discussion

19.14 Discussion

19.15 $Z = 36$, $Z = 54$

19.16 Glossary

19.17 Discussion

19.18 Discussion

Study Guide
Brief Answers

20.1 0.14 c or 4.2×10^{7} m/sec

20.2 3.7×10^{-14} newtons

20.3 $m_o v^2$ and $m_o v$

20.4 (a) changes are too small
(b) 10^{-12} kg

20.5 (a) 27×10^{32} joules
(b) 3×10^{16} kg
(c) 5×10^{-7} %
(d) Rest mass

20.6 Discussion
20.7 (a) 1.2×10^{-22} kg·m/sec
(b) 1.1×10^{-22} kg·m/sec
(c) 2.4×10^{-22} kg·m/sec
(d) 1.1×10^{-22} kg·m/sec

20.8 $p = 1.7 \times 10^{-27}$ kg·m/sec
$v = 1.9 \times 10^{3}$ m/sec

20.9 Discussion

20.10 Diagram

20.11 6.6×10^{6} m/sec

20.12 3.3×10^{-33} m

20.13 Proof

20.14 Discussion

20.15 $\Delta x = 3.3 \times 10^{-31}$ m

20.16 Discussion

20.17 Discussion

20.18 (a) 2 m, 3.3×10^{-24} kg·m/sec, 3.3×10^{-25} m/sec
(b) 0.2 m, 3.3×10^{-33} kg·m/sec, 5×10^{-8} m/sec
(c) 2×10^{-6} m, 3.3×10^{-28} kg·m/sec, 3.3×10^{-6} sec
(d) 2×10^{-10} m, 3.3×10^{-24} kg·m/sec, 3.3×10^{6} m/sec

20.19 Discussion

20.20 Discussion

20.21 Discussion

20.22 Discussion

Solutions to Chapter 17 Study Guide

17.1

The atomic mass of zinc is listed as 65.4 units and oxygen as 16.0. Therefore, the mass of a combination of 1 zinc atom and 1 oxygen atom is:

$$\underset{\text{(zinc)}}{65.4} + \underset{\text{(oxygen)}}{16.0} = \underset{\text{(zinc oxide)}}{81.4}$$

The percentage of zinc in the combination is the fraction of zinc times 100:

$$\frac{65.4}{81.4} \times 100 = 80.3\%\ \text{zinc.}$$

Therefore, the percentage of oxygen is (100 - 80.3)% = 19.7%.

17.2

As in problem 17.1, the molecular mass for the compound is computed from the atomic masses listed:

$$\underset{\text{(zinc)}}{65.4} + \underset{\text{(chlorine)}}{2(35.5)} = \underset{\text{(zinc chloride)}}{136.4}$$

The percentage of zinc is therefore

$$\frac{65.4}{136.4} \times 100 = 47.9\%.$$

17.3

Using an equation similar to (17.4) we have:

$$\frac{\text{mass of hydrogen}}{\text{mass of nitrogen}} = \frac{\text{mass of H atom}}{\text{mass of N atom}} \times \frac{\text{number of H atoms}}{\text{number of N atoms}}.$$

Since the mass of nitrogen obtained was 4.11 g, the mass of hydrogen is (5.00 - 4.11)g = 0.89 g. From the formula NH_3, the ratio of number of H atoms to number of N atoms is 3/1. Therefore

$$\frac{0.89\ \text{g}}{4.11\ \text{g}} = \frac{3}{1} \times \frac{\text{mass of H atom}}{\text{mass of N atom}}$$

or mass of N atom = 13.9 × mass of H atom. Using values of the atomic masses from the modern version of the periodic table yields a similar result:

$$\frac{\text{mass of H}}{\text{mass of N}} = \frac{1.01}{14.0} = \frac{1}{13.9}.$$

17.4

Problem 17.3 states that 5.00 g of ammonia is composed of 4.11 g of nitrogen and 0.89 g of hydrogen. This ratio will be the same for any quantity of ammonia. Therefore,

$$\frac{4.11\ \text{g nitrogen}}{5.00\ \text{g ammonia}} = \frac{x\ \text{g nitrogen}}{1200\ \text{g ammonia}}$$

$$x = 986.$$

With this quantity of nitrogen, 214 g of hydrogen will be needed to form 1200 g (1.2 kg) of ammonia.

17.5

The ratio of NH_2 would be:

$$\text{mass of N atom} = \frac{2}{1} \times \frac{4.11\ \text{g}}{0.89\ \text{g}} \times \text{mass of H atom}$$

$$= 9.23 \times \text{mass of H atom.}$$

17.6

Using an equation similar to (17.4) we have:

for NO:
$$\frac{\text{mass of N in sample}}{\text{mass of O in sample}} = \frac{1}{1} \times \frac{\text{mass of N atom}}{\text{mass of O atom}},\ \text{or}$$

$$\frac{\text{mass of N atom}}{\text{mass of O atom}} = \frac{0.47\ \text{g}}{(1-0.47)\text{g}}$$

$$= 0.89.$$

for NO_2:
$$\frac{\text{mass of N in sample}}{\text{mass of O in sample}} = \frac{1}{2} \times \frac{\text{mass of N atom}}{\text{mass of O atom}},\ \text{or}$$

$$\frac{\text{mass of N atom}}{\text{mass of O atom}} = 2 \times 0.89$$

$$= 1.78.$$

for N_2O:
$$\frac{\text{mass of N in sample}}{\text{mass of O in sample}} = \frac{2}{1} \times$$

$$\frac{\text{mass of N atom}}{\text{mass of O atom}} = \frac{1}{2} \times 0.89$$

$$= 0.445.$$

These numbers are the ratios of the atomic masses. If oxygen is defined to have an atomic mass of 16.00, then the calculated atomic mass of N would be:

(a) for NO: 16 × .89 = 14.1

(b) for NO_2: twice as much or 28.2

(c) for N_2O: half as much or 7.0.

17.7

If the value 1 is assigned to hydrogen the value for oxygen will be:

$$\text{mass of O atom} = \frac{\text{mass of O atom}}{\text{mass of H atom}} \times$$

$$\text{mass of H atom} = \frac{7}{1} \times 1 = 7.$$

And the value for nitrogen will be

$$\text{mass of N atom} = \frac{\text{mass of N atom}}{\text{mass of O atom}} \times$$

$$\text{mass of O atom} = \frac{8}{9.2} \times 7 \approx 6.$$

17.8

If we assign Cl the valence of 1, then the valences of the other elements are:

sodium:	1
calcium:	2
aluminum:	3
tin:	4
phosphorus:	5.

It should be noted, however, that the information given does not preclude the possibility of other valence numbers for the same elements.

17.9

(a) The reversals are:

Elements	Atomic Numbers	Atomic Mass
Ar and K	18 and 19	39.9 and 39.1
Co and Ni	27 and 28	58.9 and 58.7
Te and I	52 and 53	127.6 and 126.9
Th and Pa	90 and 91	232.0 and (231)
U and Np	92 and 93	238.0 and (237)
Es and Fm	99 and 100	(254) and (252)
Md and No	101 and 102	(256) and (254)

(b) If the elements are ordered according to increasing mass, then in addition to the gradual increase of mass which defines our ordering there is a periodic change of chemical and physical properties (e. g., of valence and of atomic volume). This periodicity is almost perfect except for occasional reversals (above) where positions according to mass do not correspond with the expected positions according to their properties. If the elements were ordered according to increasing atomic number (number of protons in the nucleus) these reversals would still exist, but lose their significance. The atomic numbers are never "reversed."

17.10

There is no true periodicity but there are noticeable regularities. For example, every return to zero valence (at atomic numbers 2, 10, 18, 36 and 54) is followed by a steady rise in valence toward a maximum valence of 7 (at atomic numbers such as 17, 25, 41 and 53). We note that the zero valence elements are the inert gases and that they are followed by the chemically very active elements Li, Na, K, etc. There thus appears to be a significant relationship between the valence of the atoms of an element and the chemical activity of that element. The more active elements have valences of 1 or 7. The physical significance lies in the electron structure of the atom. This will be discussed in Chapters 19 and 20.

The question of the necessity of this relationship is a philosophic one. There have appeared in nature very few coincidences which had no deeper connection than a chance correlation. Therefore, physicists tend to believe that for every correlation there is an underlying reason.

17.11

Data for both melting and boiling points are shown in the graph below. Two regularities can easily be seen. There is a trend for boiling and melting points to increase with increasing atomic mass. Superimposed on this trend is a very pronounced rise and fall of melting and boiling point temperature with the same periodicity as the Periodic Table—the lowest temperatures regularly occurring in the inert gases.

17.12

Since 96,500 coul will produce 1.008 g of hydrogen, a portion of this quantity will produce a smaller quantity of hydrogen. A simple proportion may be written as follows: (Let x represent the number of grams of hydrogen.)

$$\frac{1.008 \text{ g}}{96{,}500 \text{ coul}} = \frac{x}{3 \text{ amp} \times 3600 \text{ sec}}$$

$$3 \text{ amp} \times 3600 \text{ sec} = 10{,}800 \text{ coul}$$

$$x = \frac{1.008 \text{ g}}{96{,}500 \text{ coul}} \times 10{,}800 \text{ coul}$$

$$= 0.113 \text{ g hydrogen.}$$

In like manner, a proportion for oxygen is set up:

$$\frac{8.00 \text{ g oxygen}}{96{,}500 \text{ coul}} = \frac{x}{10{,}800 \text{ coul}}$$

$$x = 0.895 \text{ g oxygen.}$$

Note that the ratio of hydrogen to oxygen remains the same:

$$1.008/8.00 = 0.112/0.895 \approx 1/8.$$

17.13

As in problem 17.12, a proportion may be set up:

(a) $$\frac{32.69 \text{ g zinc}}{96{,}500 \text{ coul}} = \frac{x}{0.5 \text{ amp} \times 300 \text{ sec}}$$

$$x = 0.05 \text{ g zinc.}$$

(b) for 30 min:

$$x = \frac{32.69 \text{ g zinc} \times 0.5 \text{ amp} \times 1800 \text{ sec}}{96{,}500 \text{ coul}}$$

$$= 0.30 \text{ g zinc.}$$

(c) for 120 min:

$$x = \frac{32.69 \text{ g zinc} \times 0.5 \text{ amp} \times 7200 \text{ sec}}{96{,}500 \text{ coul}}$$

$$= 1.2 \text{ g zinc.}$$

17.14

(a) Table 17.4 indicates that 96,540 coul will produce 35.45 g of chlorine. If $2.0 \times 1.2 \times 10^3$ coul (i. e., 2.0 amp × 1200 sec) are used, the following proportion will indicate the amount of chlorine produced:

$$\frac{35.45 \text{ g chlorine}}{96{,}540 \text{ coul}} = \frac{x}{2.0 \text{ amp} \times 1200 \text{ sec}}$$

$$= 0.88 \text{ g chlorine.}$$

(b) Table 17.4 does not indicate the quantity of iodine that would be produced by one faraday of charge. This can be found from Faraday's second law of electrolysis. Since iodine has a valence of 1, the amount produced by 96,540 coul is 126.9 (atomic mass)/1 (valence). The proportion is written:

$$\frac{126.9 \text{ g iodine}}{96{,}540 \text{ coul}} = \frac{x}{2.0 \text{ amp} \times 1200 \text{ sec}}$$

$$x = 3.14 \text{ g iodine.}$$

(c) The quantity of zinc in part (b) would be identical to that produced in part (a) because the mass liberated is proportional to the amount of charge (Faraday's first law) which is the same in both cases.

17.15

From Faraday's investigations with electrolysis, the implication arises that electricity is composed of electrically charged discrete quantities. Observing that the amount of a substance liberated by a given quantity of electricity varies from one substance to another, Faraday concluded that this liberated amount is proportional to (1) the total quantity of electricity that passes through the electrolyte and (2) the chemical equivalent "weight" of the substance liberated. It appears then, that the released substance can accept only certain amounts of electricity and that the electricity is able to give up certain discrete amounts. Thus, electricity might be composed of "atoms" which can unite and become an integral part of "chemical atoms."

17.16

Students have been told of atoms authoritatively since their earliest grades. Most of them accept the notion on faith without ever questioning it. However, this question provides an excellent opportunity for those few "atom skeptics" you may have in your group who feel that atoms really don't exist. If you allow them to go far enough, these nonbelievers will soon be accepting the existence of atoms by means of their own astute reasoning.

The second portion of the question can be employed in demanding clear-cut evidence to reinforce and defend the faith of your "atom believers."

17.17

Modern man finds himself, of necessity, deeply involved in the many-faceted society of the twentieth century. His ties with political, religious, cultural and civic groups demand an intimate and meaningful bond and commitment to the ideals and obligations they profess. Although still an individual, man is firmly an integral part of society and must relate conscientiously to the many aspects of modern life by close ties. He must be capable of "bonding with multiplicity;" he must be "multivalent."

17.18

The idea that matter is composed of atoms was proposed by the Greek philosophers Empedocles and Democritus between 500 and 400 B. C. The Greeks in general, devised the theory principally in an "arm chair" fashion with no attempt at confirming the theory in terms of physical experimentation. The establishment of a scientific theory involves more than just a "bright idea." The idea must be able to explain known phenomena and to predict new phenomena. The Greek theory did not do this.

The experimenters of the nineteenth century, by contrast, desired to devise theories which would account for the observed properties of matter. The explanation or "mental model" exists in the mind and should reproduce the observed behavior of matter at least approximately. In addition, the nineteenth century atomic theory included a predictive function which would include, possibly, behavior not yet observed. When the model does predict as well as explain behavior, the theory is treated with more confidence.

Solutions to Chapter 18 Study Guide

18.1

(a) Where the effects of the electric and magnetic fields cancel, we have

$$qE = qvB, \quad \text{or} \quad v = \frac{E}{B}, \text{ and}$$

$$\text{since } E = \frac{V}{d}, \quad v = \frac{V}{Bd};$$

$$\text{so } v = \frac{200 \text{ volts}}{1.0 \times 10^{-3} \dfrac{\text{N}}{\text{amp}\cdot\cancel{\text{m}}} \times 0.01\ \cancel{\text{m}}}$$

$$= 2.0 \times 10^{7} \text{ m/sec}.$$

$$\left(\frac{\text{volt}\cdot\text{amp}}{\text{N}} = \frac{\dfrac{\cancel{\text{N}}\cdot\text{m}}{\cancel{\text{coul}}} \cdot \dfrac{\cancel{\text{coul}}}{\text{sec}}}{\cancel{\text{N}}} = \frac{\text{m}}{\text{sec}}\right)$$

(b) When the magnetic field acts alone, a circular orbit results, and

$$qvB = \frac{mv^2}{R}, \quad \text{or} \quad \frac{q}{m} = \frac{v}{BR}$$

$$\frac{q}{m} = \frac{= 2.0 \times 10^{7} \dfrac{\text{m}}{\text{sec}}}{1.0 \times 10^{-3} \dfrac{\text{N}}{\text{amp}\cdot\cancel{\text{m}}} \times 0.114\ \cancel{\text{m}}}$$

$$= 1.8 \times 10^{11} \text{coul/kg}$$

$$\left(\frac{\text{amp}\cdot\text{m}}{\text{N}\cdot\text{sec}} = \frac{\dfrac{\text{coul}}{\cancel{\text{sec}}} \cdot \cancel{\text{m}}}{\dfrac{\text{kg}\cdot\cancel{\text{m}}}{\cancel{\text{sec}^2}} \cdot \cancel{\text{sec}}} = \frac{\text{coul}}{\text{kg}}\right)$$

18.2

Since 1 amp = 1 coul/sec past a given point, we need to show how many basic units of charge (electron or proton charges) are equal to one coulomb (call n this number of electrons). Then n × charge on each electron = 1 coulomb

$$\text{or} \quad n = \frac{1 \text{ coul}}{1.6 \times 10^{-19} \text{ coul/electron}}$$

$$\text{so} \quad n = 6.25 \times 10^{18} \text{ electrons}.$$

18.3

The final energy = the initial energy. Consider the simple case of a uniform electrostatic field between cathode and anode. Then the force on the electron is $q_e E$ everywhere. The electron leaves the plate with some kinetic energy. It is slowed by the electrostatic field and stops a distance d from the cathode. The initial kinetic energy must equal the work done by the electron so

$$q_e Ed = \tfrac{1}{2}mv_I^{\,2}.$$

Now the electron is pushed back toward the cathode by the field. When it hits it has travelled the same distance d back. The work done on the electron is equal to the kinetic energy it gains so

$$q_e Ed = \tfrac{1}{2}mv_F^{\,2}.$$

We see that $\tfrac{1}{2}mv_I^{\,2} = q_e Ed = \tfrac{1}{2}mv_F^{\,2}$. The initial and final kinetic energies are equal.

Similar arguments can be made by considering the potential energy-kinetic energy interplay, or by considering an arbitrary force which depends only on the distance from the cathode.

Some may see intuitively that the energies are equal from consideration of the symmetry of the electron path. Symmetry arguments are occasionally misleading though, and should, where possible, be flushed out with an argument like the one above.

18.4

The light energy is either absorbed by the crystal lattice as a whole, increasing its thermal energy, or is reflected.

18.5

The work function $W = hf_o$,

$$\text{so } f_o = \frac{W}{h} = \frac{10^{-19} \text{ joule}}{6.6 \times 10^{-34} \text{ joule}\cdot\text{sec}}$$

$$= 1.5 \times 10^{14} \text{ cycles/sec}.$$

This corresponds to a wavelength given by

$$\lambda = \frac{3 \times 10^{8} \text{ m/sec}}{1.5 \times 10^{14} \text{ cycles/sec}}$$

$$\approx 2 \times 10^{-6} \text{ m} \quad \text{or} \quad 2000\text{Å}.$$

This wavelength lies in the ultra-violet region of the spectrum.

18.6

The energy of a photon is given by

$$E = hf, \text{ and since } f = \frac{c}{\lambda}$$

$$E = \frac{hc}{\lambda} = \frac{6.6 \times 10^{-34} \text{ joule}\cdot\text{sec} \times 3 \times 10^{8} \text{ m/sec}}{5 \times 10^{-7} \text{ m}}$$

$$= 4 \times 10^{-19} \text{ joule}.$$

For $\lambda = 5 \times 10^{-8}$ m, $E = 4 \times 10^{-18}$ joule.

18.7

Light of threshold frequency has an energy which is just sufficient to free an electron from the metal. The energy associated with that minimum frequency (f_o) is called the work function (W) of the metal, and $W = hf_o$. For copper,

$$W = 6.6 \times 10^{-34} \text{ joule·sec} \times 1.1 \times 10^{15} \frac{\text{cycles}}{\text{sec}}$$

$$= 7.3 \times 10^{-19} \text{ joule.}$$

When light of greater than threshhold frequency is used, the photoelectrons will be emitted with a maximum kinetic energy given by

$$KE_{max} = hf - W$$

$$= 6.6 \times 10^{-34} \text{ joule·sec} \times 1.5 \times 10^{15} \frac{\text{cycles}}{\text{sec}}$$

$$= 7.3 \times 10^{-19} \text{ joule.}$$

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$$= 6.6 \times 10^{-34} \text{ joule·sec} \times 1.5 \times 10^{15} \frac{\text{cycles}}{\text{sec}}$$

$$- 7.3 \times 10^{-19} \text{ joule,}$$

$$= (9.9 - 7.3) \times 10^{-19} \text{ joule}$$

$$= 2.6 \times 10^{-19} \text{ joule.}$$

(Alternatively, $KE_{max} = hf - hf_o = h(f - f_o)$, etc.)

Since 1 eV = 1.6×10^{-19} joule,

$$KE_{max} = \frac{2.6 \times 10^{-19} \text{ joule}}{1.6 \times 10^{-19} \text{ joule/eV}} = 1.6 \text{ eV.}$$

18.8

The energy of a photon which will cause the emission is given by

$$f_o = \frac{W}{h} = \frac{2.0 \text{ eV} \times 1.6 \times 10^{-19} \text{ joule/eV}}{6.6 \times 10^{-34} \text{ joule·sec}}$$

$$= 4.9 \times 10^{14} \text{cycles/sec}$$

18.9

(a) For each photon, $E = \frac{hc}{\lambda}$

$$= \frac{6.6 \times 10^{-34} \text{ joule·sec} \times 3.0 \times 10^{8} \text{ m/sec}}{5.0 \times 10^{-7} \text{ m}}$$

$$= 4.0 \times 10^{-19} \text{ joule.}$$

Thus, the number of photons required to cause the given intensity would be

$$= \frac{100 \text{ joules}}{4.0 \times 10^{-19} \text{ joules/photon}} = 2.5 \times 10^{20} \text{ photons.}$$

(b) If the atomic diameter = 1 Å, 10^{10} atoms will fit along a 1 m line and in 1 m^2 there will be $10^{10} \times 10^{10}$ atoms, or 10^{20} atoms.

Then in one second, 2.5×10^{20} photons will fall on 10^{20} atoms. If all photons are absorbed by the surface layer of atoms, each atom will absorb on the average 2.5 photons/sec.

(c) For an average of one photon/atom it would take $\frac{1}{2.5}$ the time it takes for 2.5 photons, or 0.4 sec.

(d) The number of photons arriving per atom in 10^{-10} sec will be

$$= 2.5 \text{ photons/sec} \times 10^{-10} \text{ sec}$$

$$= 2.5 \times 10^{-10} \text{ sec.}$$

(e) The cathode area is $(0.05\text{m})^2$ or 2.5×10^{-3} m^2. Thus, the rate of arrival of photons at the cathode

$$= 2.5 \times 10^{20} \frac{\text{photons}}{m^2 \text{·sec}} \times 2.5 \times 10^{-3} \text{ m}^2$$

$$= 6.3 \times 10^{17} \text{ photons/sec.}$$

By assumption this yields 6.3×10^{17} electrons/sec. The current is thus

$$= 6.3 \times 10^{17} \frac{\text{electrons}}{\text{sec}} \times 1.6 \times 10^{-19} \frac{\text{coul}}{\text{electron}}$$

$$= 0.1 \text{ coul/sec,} \quad \text{or} \quad 0.1 \text{ amp.}$$

(Note: you may object that the above answers are given to more significant figures than are warranted by the problem as stated. You would be correct. Significant figures are very important when making calculations from data; but the round numbers given in this problem are contrived for the purpose of an exercise, to keep arithmetic from getting in the way of ideas.)

18.10

By definition, 1 watt = 1 joule/sec. However, the energy transformed into light is only 5 per cent of this, or 0.05 joule. For one photon, E = hf, and for n photons, E = nhf. Thus substituting

$$f = \frac{c}{\lambda}, \quad E = \frac{nhc}{\lambda} \quad \text{or} \quad n = \frac{E\lambda}{hc};$$

$$n = \frac{0.05 \text{ joules} \times 5 \times 10^{-7} \text{ m}}{6.6 \times 10^{-34} \text{ joule·sec} \times 3 \times 10^{8} \frac{\text{m}}{\text{sec}}}$$

$$= 1.3 \times 10^{17} \text{photons.}$$

18.11

We are given $hf_{max} = q_e V$, thus

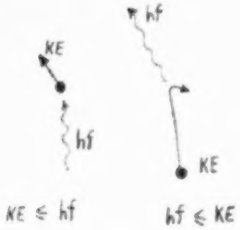
$$f_{max} = \frac{q_e V}{h} = \frac{1.6 \times 10^{-19} \text{ coul} \times 5 \times 10^{4} \text{ volts}}{6.6 \times 10^{-34} \text{ joule·sec}}$$

$$= 1.2 \times 10^{-19} \text{ cycles/sec.}$$

$$\left(\frac{\text{coul}\cdot\text{volt}}{\text{joule}\cdot\text{sec}} = \frac{\cancel{\text{coul}}\cdot\frac{\cancel{\text{joule}}}{\cancel{\text{coul}}}}{\cancel{\text{joule}}\cdot\text{sec}} = \text{cycles/sec} \right)$$

18.12

In the one case, electrons are given kinetic energy by photons. In the other, photons are produced when an electron loses kinetic energy. Since in both cases the energy of the second comes from the first, the second can not have more energy than the first.



18.13

The energy of a photon is given by $E = \frac{hc}{\lambda}$. This energy can at most be as large as the electron energy $q_e V$. Thus, for minimum wavelength (max. energy)

$$q_e V = \frac{hc}{\lambda_{min}}, \qquad V = \frac{hc}{q_e \lambda_{min}}$$

$$V = \frac{6.6 \times 10^{-34} \text{ joule}\cdot\text{sec} \times 3 \times 10^{8} \text{ m/sec}}{1.6 \times 10^{-19} \text{ coul} \times 10^{-11} \text{ m}}$$

$$= 1.2 \times 10^{5} \text{ joules/coul}, \quad \text{or}$$

$$V = 1.2 \times 10^{5} \text{ volts.}$$

This corresponds to a maximum energy of 1.2×10^{5} eV, or

$$1.6 \times 10^{-19} \text{ coul} \times 1.2 \times 10^{5} \text{ joules/coul}$$

$$= 1.9 \times 10^{-14} \text{ joules.}$$

18.14

photon: a discrete quantity of electromagnetic radiant energy; a "bundle of energy" whose value hf is proportional to the frequency of radiation.

quantum: one of the very small increments or quantities into which forms of energy are found to be subdivided.

cathode rays: emanations from the cathode electrode of a vacuum tube under the influence of an electric field; found to be electrons.

photoelectron: an electron released, generally from a metal, by means of energy absorbed from photons of light (usually ultraviolet) shining on the surface of the material.

photoelectric effect: the release of electrons from matter when exposed to certain frequencies of electromagnetic radiation.

quantum theory: that branch of modern physics based on the concept of the subdivision of electromagnetic energy into discrete quanta.

threshold frequency: the frequency of incident radiation below which the photoelectric effect will not take place.

stopping voltage: the voltage between cathode and anode in a photoelectric tube which will just stop the most energetic electrons emitted from the cathode.

classical physics: the physical theories concerning the nature of the universe and their philosophical implications which were developed prior to the advent of quantum theory.

x rays: electromagnetic radiation of short wavelength produced by electron bombardment of matter.

18.15

(a) Einstein's interpretation was not anticipated by Newton. This statement can be supported in at least three ways: (1) by considing the range and type of problem for which the models were proposed; (2) by considering qualitative differences between the models; (3) by considering the precision of the models in predicting experimental results.

(b) As indicated in part (a), Newton's model was qualitative and tied to classical particle mechanics. Newton could not have predicted the slope, intercept, or general form of the energy vs. frequency curves. He might have had difficulty explaining the rapid emission of photoelectrons. [In a classical elastic collision between a very small particle (light) and a very large particle (electron), the large particle receives very little energy.] Newton's particles would produce results that might be expected from a classical wave picture of light.

(1) Newton was writing at a time when some of the basic qualitative features of light phenomena were being discovered and interpreted. A particle or a wave model of light could explain the light phenomena known at the time, though, on balance, Newton's particle model seemed to handle the phenomena most simply. Newton's particle model, though, was intended to explain all light phenomena. Einstein's model of light quanta was constructed at a time when the great bulk of light phenomena had been successfully explained using a wave theory of light. Hertz's experiments, in which the photoelectric effect had been discovered, were among the crowning successes of a wave theory of electromagnetic phenomena. Einstein proposed

a view of light emission and absorption which was fundamentally different from the dominant way of conceiving of light phenomena at his time. In addition, Einstein proposed light quanta as a way of explaining only a limited set of light phenomena. He conceded the essential role of a wave theory in explaining interference and other related light phenomena.

(2) Newton's light particles were distinct from particles of ordinary matter, yet similar to them. They were smaller than particles of matter and perhaps different in substance. However, their motion was to be described by the laws of classical mechanics. There is thus no reason to assume that Newton's light particles should give up all their energy when interacting with particles of "gross matter" (of course energy was not a fundamental concept for Newton).

Einstein's quantum was simply "localized energy." The interaction of this energy with particles of matter was not presumed by Einstein to be described by classical mechanics. In particular, the assumption of an all-or-nothing energy transfer was fundamental in Einstein's model and cannot be derived from classical particle-particle interactions. Einstein's equation for the quantum of energy, $E = hf$, implicitly referred to a wave model of light since frequency (f) is a wave characteristic. Newton's model could not contain such a reference to a wave characteristic.

(3) Einstein's model provided a basis for precise prediction of experimental data. Newton's model, in the main, did not. This, by itself, is enough to give credit to Einstein. Even if the models were comparable in all details, Einstein would get credit since he indicated how the model could be tested. When Newton and Hooke quarreled over who should be given credit for discovery of the inverse square law of gravitational force, Newton insisted that the individual who worked out the implications for experiment (Newton) should receive credit even though another individual (Hooke) had earlier and informally voiced the idea.

19.1

a) One way that has been used to tell if Fraunhofer lines are caused by absorption in the earth's atmosphere is to make observations over a range of zenith angles (angles measured between the line of sight and the vertical). As observations are made at larger and larger zenith angles, the light from the sun traverses a greater and greater thickness of the earth's atmosphere; this results in the darkening of absorption lines due to the earth's atmosphere.

When an absorption line is due to some element in both the sun and earth's atmospheres, the respective contributions can be found by observing reflected sunlight from the planet Mercury when it is moving toward the earth and then when moving away (i.e. at both quadratures). The sun's contribution to the total line will then be doppler-shifted to the left and to the right of the earth's contribution. (Mercury is used because of its lack of atmosphere and high speed.)

A more expensive method that can now be used is to place a spectrograph on a satellite orbiting the earth above the earth's atmosphere. Any record made by it of the sun's atmosphere would then involve only solar absorption. The Orbiting Solar Observatory, OSO-4, shown on page 124 of Unit 4, made such observations over a narrow range of wavelengths (about 1500 Å, mostly in the ultraviolet).

b) If light from the moon or planets were found to have absorption lines characteristic of sunlight, this would be strong evidence that the light was from the sun. Additional selective absorption from sunlight would indicate both the extent and composition of any atmosphere the moon and planets might have. For example, methane, which is not present on the sun, is present on Jupiter and thus causes additional absorption lines in sunlight reflected from Jupiter.

That the moon has no atmosphere (less than 10^{-6} of earth's) can be determined from its sharp occultation of background stars. Also, a lunar atmosphere would cause scattering of sunlight which would result in the "horns" of the moon appearing slightly larger. All the planets except Mercury (and perhaps Pluto) have atmospheres.

19.2

The text gives five series, but in theory there are an infinite number—one for each of the infinite number of possible values of n_f. (Most of them comprise "lines" in the radio region of the spectrum.) Only four lines—the first four of the Balmer series—are in the visible region.

19.3

The empirical formula giving the wavelength of the spectral lines of hydrogen is

$$\frac{1}{\lambda} = R_H \left[\frac{1}{n_f^2} - \frac{1}{n_i^2} \right],$$

where $R_H = 1.10 \times 10^7/\text{m}$.

For the Balmer series, $n_f = 2$. Then for $n_i = 8$, the quantity in brackets becomes

$$\frac{1}{2^2} - \frac{1}{8^2}, \text{ or } \frac{1}{4} - \frac{1}{64} = \frac{15}{64};$$

$$\text{for } n_i = 10, \frac{1}{2^2} - \frac{1}{10^2}, \text{ or } \frac{24}{100};$$

$$\text{for } n_i = 12, \frac{1}{2^2} - \frac{1}{12^2}, \text{ or } \frac{35}{144}.$$

Solving the above formula for λ, and evaluating for

$$n_i = 8, \lambda = \frac{1 \text{ m}}{1.10 \times 10^7 \times \frac{15}{64}} = 3.88 \times 10^{-7} \text{ m or } 3880 \text{ Å};$$

$$n_i = 10, \lambda = \frac{1 \text{ m}}{1.10 \times 10^7 \times \frac{24}{100}} = 3.79 \times 10^{-7} \text{ m or } 3790 \text{ Å};$$

$$n_i = 12, \lambda = \frac{1 \text{ m}}{1.10 \times 10^7 \times \frac{35}{144}} = 3.74 \times 10^{-7} \text{ m or } 3740 \text{ Å}.$$

In Table 19.1, where n increases from 3 to 6, and here, where n increases from 8 to 12, it is apparent that λ decreases as n increases. Also, the increment by which λ decreases as n increases, say, from 6 to 8, 8 to 10, 10 to 12, becomes smaller and smaller, implying that as higher quantum numbers are approached, the difference between energy states approaches zero. This is precisely what happens, as can be seen from Fig. 19.11.

This problem illustrates a very important trend, namely, as high quantum numbers are approached, quantum effects become less apparent. (In other words, the energy states get closer and closer together, until the state defined by quantum number n cannot be distinguished from the state defined by quantum number n ± 1.) This fact forms the basis of a principle which relates classical and quantum physics. This principle, first articulated by Bohr, is called the correspondence principle and says that in the limit of large quantum numbers, quantum physics merges into classical physics.

19.4

a) In Fig. 19.5 the bunching occurs at the high frequency end of the spectrum. Fig. 19.11 shows that the high frequency transitions involve large quantum numbers. As seen in the preceding problem solution, high values of n lead to nearly identical values of λ, so the bunching is predicted by the formula.

b) $n_i = \infty$

c) The Lyman series has $n_f = 1$, so the series limit, where $n_i = \infty$, is

$$\lambda = \frac{1\ \text{m}}{1.10 \times 10^7 \times \left(\frac{1}{1^2}\right)} = 0.910 \times 10^{-7}\ \text{m} \text{ or } 910\ \text{Å};$$

The Balmer series has $n_f = 2$, so the series limit is

$$\lambda = \frac{1\ \text{m}}{1.10 \times 10^7 \times \left(\frac{1}{2^2}\right)} = 3.64 \times 10^{-7}\ \text{m} \text{ or } 3640\ \text{Å};$$

The Paschen series has $n_f = 3$, so the series limit is

$$\lambda = \frac{1\ \text{m}}{1.10 \times 10^7 \times \left(\frac{1}{3^2}\right)} = 8.18 \times 10^{-7}\ \text{m} \text{ or } 8180\ \text{Å}.$$

d) The series limit is 910 Å for the Lyman series. This wavelength corresponds to an energy given by $\frac{hc}{\lambda}$;

$$E = \frac{6.6 \times 10^{-34}\ \text{joule}\cdot\text{sec} \times 3.0 \times 10^8\ \text{m/sec}}{0.910 \times 10^{-7}\ \text{m}}$$

$$= 21.8 \times 10^{-19}\ \text{joule, or}$$

$$\frac{21.8 \times 10^{-19}\ \text{joule}}{1.6 \times 10^{-19}\ \text{joule/eV}} = 13.6\ \text{eV}.$$

19.5

The Thomson and Rutherford models of the atom are similar in the following ways: atoms contain positive and negative charge in equal amounts; nearly all of the mass of the atom is associated with the positive charge; the diameter of the atom is of the order of 10^{-10} m.

The models differ in that in the Thomson atom the positive charge is spread out through the entire volume of the atom; in the Rutherford model the positive charge is concentrated (localized) into a very small volume at the center of the atom. Also, in the Thomson model, the electrons are distributed throughout the positive charge; in the Rutherford model the electrons are separated from the positive charge, being distributed around the positive charge in some undefined way. Lastly, Rutherford's model has much empty space; Thomson's atom is "full" of charged matter.

19.6

a) Lenard's model is like those of Thomson and Rutherford in having positive and negative electricity in equal amounts so that the atom as a whole is neutral. It is similar to Rutherford's model in having much empty space. But Lenard's model did not distinguish between the electrons and the positive charge as did the Thomson and Rutherford models. Hence, in Lenard's model, the positive charge was not associated with nearly all the mass of the atom.

b) No. It would not be possible to account for the back-scattering of swift α particles. A neutral dynamide would have to have a mass no greater than that of a hydrogen atom (if we assume that a hydrogen atom consists of one dynamide). Hence a heavier nucleus, such as those used by Rutherford in his scattering experiments, would contain many dynamides, each with much smaller mass than an α particle. A collision between an α particle and a neutral dynamide would be like a billiard ball collision of a moving heavy ball with a lighter, stationary ball. The heavier ball would not be deflected significantly from its forward direction in a single collision. The angular distribution of the scattered α particles would then be in a small angle about the direction of the incident (α particle) beam, with no backward scattering.

19.7

Certainly the burden of proof is on the author. The nuclear model of the atom is consistent with the scattering experiments, it has served to suggest new experiments, and in general it nicely ties together data pertaining to atomic structure.

Furthermore the author in question proposes that the atom is a small neutral particle. If this were the case, how would he account for the scattering forces that cause the observed scattering? In the nuclear model, scattering forces between a positively charged α particle and positively charged nucleus are Coulomb forces; what kind of forces would act between an alpha particle and a small neutral atom? It is not at all clear what the origin of the required forces would be; hence it would be very difficult to account for the observed alpha scattering.

19.8

One Angstrom unit is 10^{-8} cm. Hence:

$$\frac{d_{\text{nucleus}}}{d_{\text{atom}}} = 10^{-4}$$

It might be interesting to point out that the density of the nucleus can be estimated.

$$\frac{\text{volume of nucleus}}{\text{volume of atom}} = \left[\frac{d_n}{d_a}\right]^3 = 10^{-12}.$$

This result indicates that the density of the nucleus must be enormous. The mass of a light atom is of the order of 10^{-24} g. Thus, taking the volume of the atom as approximately $(10^{-12}\text{ cm})^3$, the density

$$= \frac{10^{-24}\text{ g}}{10^{-36}\text{ cm}^3} = 10^{12}\text{ g/cm}^3.$$

19.9

The magnification is $\dfrac{10^{-2}\text{ cm}}{1.5 \times 10^{-13}\text{cm}}$, or 6.7×10^{11}. Thus, the magnified radius of the first Bohr orbit would be the product of the magnification and the actual radius (given on page 82); $6.7 \times 10^{11} \times 5.29 \times 10^{-9}$ cm

$= 3.5 \times 10^2$ cm or 3.5 meters.

19.10

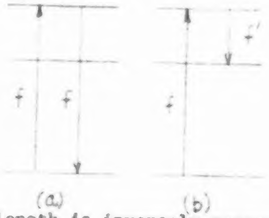
Open-ended question, but the following comment is pertinent: Although at the beginning of this century the atom may have been considered an artificial idea, introduced to explain a limited set of phenomena, the vast variety of experimental evidence which has proved to be consistent with our idea of atoms makes the atom just as real as, say, Jupiter.

19.11

The Bohr model can account for the lines of absorption spectra if it is assumed that the orbital electron can absorb a light quantum only if the energy so absorbed raises the electron into another allowed orbit. The absorption of light is then the exact inverse of emission and every absorption line should correspond exactly to an emission line, in agreement with experiment.

19.12

When a substance is illuminated with ultraviolet light of frequency f, an atom of the substance in its ground state absorbs energy in the amount hf. The atom is raised from its ground state to an excited state; or, to put it another way, an electron is raised from its normal orbit into an outer orbit. The electron can then drop back to some lower orbit between its initial orbit and the one to which it is raised. In other words, it can give up either the energy hf [see (a) below] or an amount of energy hf' [see (b) below]. It is seen that $hf' < hf$ or $f' < f$.



The wavelength is inversely proportional to frequency. Hence $\lambda' > \lambda$, which is what Stokes found. Ultraviolet light includes frequencies greater than those of visible light. Thus, in fluorescence caused by ultraviolet light, much of the light reradiated is visible, that is, f is ultraviolet while f' is visible.

19.13

a) The concept of atoms as expressed by Newton is quite similar to that attributed to Leucippus and Democritus. One difference is that, according to Leucippus and Democritus, atoms are eternal; according to Newton, God created atoms in the beginning. For Leucippus and Democritus all atoms are of the same kind, but differ in size, shape and position; Newton's atoms have sizes, shapes, and "such other properties..., as most conduced to the End for which he formed Them." There is a theological aspect to Newton's views which is not found in the Greek atomists.

b) Dalton hypothesizes that each element (an idea not mentioned by Newton or Leucippus and Democritus) consists of a characteristic kind of identical atoms: the atoms of an element "are perfectly alike in weight and figure, etc." Between Newton and Dalton much progress had been made in the understanding of the concepts of chemical element and chemical compound. Dalton could also, therefore, make an hypothesis concerning the details of the formation of compounds by the atoms of different elements.

c) In the Rutherford-Bohr model, the atom was no longer "solid," "impenetrable," "uncuttable," or "indivisible." The atom of Rutherford and Bohr consists of a nucleus and electrons and empty space. In view of what was known about radioactivity in 1913, especially since it was known that atoms could emit α particles, Rutherford and Bohr made no detailed hypotheses about the nature of the nucleus. It was already evident that the nucleus (the only place α particles could come from) was not "indivisible."

19.14

An atom normally has its electrons in the lowest possible energy states. An atom of potassium has Z = 19 and thus has 19 electrons which will be situated as follows: (see chart on page 91, noting that each circle represents a pair of electrons) 2 electrons in K shell, 8 in L, 8 in M, then the one electron remaining would be in the lowest energy level of the N shell, because that level is lower in energy than the five pairs of locations still open in the M shell.

19.15

Refer to the chart on page 91. Starting with argon ($Z = 18$), we continue adding electrons in pairs; 2 electrons in the N shell, 10 in M, 6 more in N—now we have a stable arrangement of 8 outer electrons. The element having this number of electrons is krypton, $Z = 36$.

To find the next inert gas after krypton, we continue from $Z = 36$, adding 2 electrons in the O shell, 10 in N, 6 more in O shell—now we have another stable configuration of 8 electrons in the outer shell. The element having this number of electrons is xenon, $Z = 54$.

19.16

Glossary of some of the terms which should be defined are:

- α particles
- Bohr orbit
- empirical relation
- energy-level diagram
- excitation energy
- ground state
- line absorption spectrum
- line emission spectrum
- nuclear atom
- nucleus
- optical spectra
- planetary atom
- quantum mechanics
- radioactive substances
- scintillation
- shell
- spectrum
- stationary states
- x-ray spectra

19.17

This is open-ended, but we would like to point out the following: In particle dynamics, collisions can be considered singly, and the collision of a collection of particles is simply the summation of individual collisions. It appears, however, that a different level of analysis is needed for mental phenomena. Thought comprises the interaction of complex systems, rather than of discrete entities. The model is more like a committee meeting than a game of billiards. Furthermore, the particles of atomism are of a limited number of types; those of the same type being in fact identical. However, brain cells and the system of connections among them are infinitely variable.

19.18

a) The statement in the text really deals with all the quantized things the student would know. Additional things they wouldn't know are "strangeness" and "baryon number".

b) Some properties or things outside physics that can be thought of as being quantized are the following:

salary increases in large corporations
consumer prices (in units of 1 cent)
snow, rain, hail, sleet (fortunately!)
formal education (in units of courses)
letter grades (not per-cent grades, though)
dates ("tomorrow" suddenly becomes "today," etc.)

Note: we do not mention quantities which are intrinsically collections of things.

Solutions to Chapter 20 Study Guide

20.1

The mass m at relativistic speed v is given by

$$m = \frac{m_o}{\sqrt{1 - \frac{v^2}{c^2}}}, \text{ so } \frac{m}{m_o} = \frac{1}{\sqrt{1 - \frac{v^2}{c^2}}}.$$

In this problem, $m = 1.01\, m_o$, so $\frac{1.01\, m_o}{m_o} = \frac{1}{\sqrt{1 - \frac{v^2}{c^2}}}$

Squaring both sides, $1.02 = \frac{1}{1 - \frac{v^2}{c^2}}$,

or $1 - \frac{v^2}{c^2} = \frac{1}{1.02} = 0.98.$

Thus, $\frac{v^2}{c^2} \doteq 0.02$, $\frac{v}{c} = 0.14$,

$v = 0.14\, c$, or $0.14 \times 3.0 \times 10^8$ m/sec

$= 4.2 \times 10^7$ m/sec.

20.2

The relativistic mass (m) is related to the rest mass (m_o) by

$$m = \frac{m_o}{\sqrt{1 - \frac{v^2}{c^2}}} = \frac{9.1 \times 10^{-31}\text{ kg}}{\sqrt{1 - \left(\frac{0.6c}{c}\right)^2}}$$

$$= \frac{9.1 \times 10^{-31}\text{ kg}}{\sqrt{1 - 0.36}} = \frac{9.1 \times 10^{-31}\text{ kg}}{0.80}$$

$$= 11.4 \times 10^{-31}\text{ kg}.$$

The centripetal force required is mv^2/R, but

$v = 0.60\, c$, so

$$F = \frac{m(0.60c)^2}{R}$$

$$= \frac{11.4 \times 10^{-31}\text{ kg} \times (0.60 \times 3 \times 10^8\text{ m/sec})^2}{1.0\text{ m}}$$

$$= 11.4 \times 10^{-31}\ (3.24 \times 10^{16})\text{ kg·m/sec}^2$$

$$= 3.7 \times 10^{-14}\text{ newtons}.$$

20.3

a) By substitution of $v/c = 0.1$, the series becomes $1 + 1/2\ (0.1)^2 + 3/8\ (0.1)^4 + \ldots$

$= 1 + 0.005 + 0.0000375 + \ldots$

b) From a) only the first two terms may be necessary. The relativistic momentum is then

$$p = m_o \left(1 + 1/2\, \frac{v^2}{c^2}\right)v \text{ and since } \frac{v^2}{c^2} < 10^{-10}$$

we can neglect even that term. So, for man-sized objects, $p = m_o v$.

Similarly, the relativistic kinetic energy is

$$KE = m_o \left(1 + 1/2\, \frac{v^2}{c^2}\right)c^2 - m_o c^2 = 1/2\, m_o v^2.$$

Here the second term must be retained because the $m_o c^2$ terms drop out of the equation. Therefore, for man-sized objects, $KE = 1/2\, m_o v^2$.

20.4

a) The mass changes that are due to energy changes in chemical reactions are too small to be detected. Support for this statement is given in part b).

b) The mass change is related to the energy change as follows:

$$\Delta m = \frac{\Delta E}{c^2} = \frac{10^5\text{ joules}}{(3 \times 10^8\text{ m/sec})^2}$$

$= 1.1 \times 10^{-12}$ kg. This is only one billionth of a gram and is of course not detectable in chemical reactions.

$$\left(\frac{\text{joules}}{\text{m}^2/\text{sec}^2} = \text{N}\cdot\cancel{\text{m}}\cdot\frac{\text{sec}^2}{\text{m}^{\cancel{2}}} = \frac{\text{kg}\cdot\cancel{\text{m}}}{\cancel{\text{sec}^2}}\cdot\frac{\cancel{\text{sec}^2}}{\cancel{\text{m}}} = \text{kg}\right)$$

20.5

a) The kinetic energy of the earth is

$$KE = 1/2\, m_o v^2 = 1/2 \times 6.0 \times 10^{24}\text{ kg }(3 \times 10^4\text{ m/sec})^2$$

$$= 27 \times 10^{32}\text{ joules}.$$

b) The mass equivalent is $\Delta m = \frac{\Delta E}{c^2}$

$$= \frac{27 \times 10^{32}\text{ joules}}{(3.0 \times 10^8\text{ m/sec})^2} = 3.0 \times 10^{16}\text{ kg}.$$

c) The percentage increase in mass is then $\frac{3.0 \times 10^{16}\text{ kg}}{6.0 \times 10^{24}\text{ kg}} \times 100\% = 5 \times 10^{-7}\%$.

d) Any measurements of the mass of the earth made on the earth will yield the rest mass of the earth, since the observers are at rest with respect to the earth.

20.6

Squire's verse has several possible meanings. First, it seems to suggest that the order and regularity suggested by the laws of Newton has been done away with by Einstein. This is not true. Einstein's work has extended our experience into new domains (speeds comparable to the speed of light). It is presumption for anyone to assume that a theory acknowledged as sound in one domain of experience can be extrapolated unchanged into another domain of experience.

Second, it implies an often-heard sentiment whenever a new theory alters the status of a former theory; namely, that scientists are always changing their minds and hence their theories should not be trusted. Such a view shows a lack of understanding of science. The best theories are those which put themselves in the greatest jeopardy. That is, a good theory is extremely suggestive in regards to new experiments any of which could change the theory. Thus, as man's abilities to experiment improve and broaden in scope, theories often have to be modified to account for the new phenomena. Remember, however, it was probably the theory itself which brought its own demise.

20.7

a) The relativistic momentum is the product of the relativistic mass and the velocity of the electron:

$$p = \frac{m_o v}{\sqrt{1 - \frac{v^2}{c^2}}} = \frac{9.1 \times 10^{-31}\ \text{kg} \times 0.4 \times 3 \times 10^{8}\ \text{m/sec}}{\sqrt{1 - \left(\frac{0.4c}{c}\right)^2}}$$

$$= \frac{9.1 \times 0.4 \times 3 \times 10^{-23}}{0.92}\ \text{kg}\cdot\text{m/sec}$$

$$= 1.2 \times 10^{-22}\ \text{kg}\cdot\text{m/sec}.$$

b) The Newtonian momentum at that speed

$$= m_o v = 9.1 \times 10^{-31}\ \text{kg} \times 0.4 \times 3 \times 10^{8}\ \text{m/sec}$$

$$= 1.1 \times 10^{-22}\ \text{kg}\cdot\text{m/sec}.$$

c) The relativistic momentum at $v = 0.8c$

$$= \frac{9.1 \times 10^{-31}\ \text{kg} \times 0.8 \times 3 \times 10^{8}\ \text{m/sec}}{\sqrt{1 - (0.8)^2}}$$

$$= 3.6 \times 10^{-22}\ \text{kg}\cdot\text{m/sec}.$$

Therefore the change in relativistic momentum due to v increasing from 0.4c to 0.8c is

$(3.6 - 1.2) \times 10^{-22}$ kg·m/sec, or

2.4×10^{-22} kg·m/sec.

d) Since the Newtonian momentum at a speed of 0.8c is simply twice its value at 0.4c, the change in Newtonian momentum is

1.1×10^{-22} kg·m/sec.

20.8

The momentum of the photon is given by

$$p = \frac{h}{\lambda} = \frac{6.6 \times 10^{-34}\ \text{joule}\cdot\text{sec}}{4 \times 10^{-7}\ \text{m}}$$

$$= 1.7 \times 10^{-27}\ \text{kg}\cdot\text{m/sec}$$

$$\left(\frac{\text{joule}\cdot\text{sec}}{\text{m}} = \frac{\text{kg}\cdot\cancel{\text{m}}\cdot\cancel{\text{sec}}}{\text{sec}^{\cancel{2}}\cdot\cancel{\text{m}}} = \text{kg}\cdot\text{m/sec}\right)$$

For an electron to have the above momentum it must have a speed given by

$$v = \frac{p}{m} = \frac{1.7 \times 10^{-27}\ \text{kg}\cdot\text{m/sec}}{9.1 \times 10^{-31}\ \text{kg}}$$

$$= 1.9 \times 10^{3}\ \text{m/sec}.$$

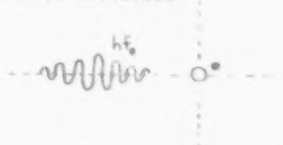
20.9

The experiments that led to the acceptance of the wave theory of light involving reflection, refraction, diffraction, interference were done in the eighteenth and early nineteenth centuries with relatively simple experimental equipment.

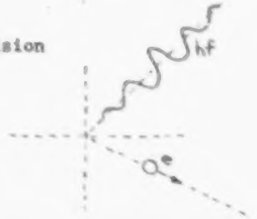
The experiments that could be interpreted only in terms of the particle aspect were the photoelectric and Compton effects. The electron had to be discovered (Thomson, 1897) and methods developed for making quantitative experiments with electrons and x rays before those effects could be analyzed and interpreted. These things were not done until the end of the nineteenth century and the first quarter of the twentieth century.

20.10

before collision



after collision



Note that the frequency of the photon has decreased. Since the energy of the photon is proportional to its frequency, the figure indicates that the photon has lost energy in the collision.

20.11

The momentum of an electron is given by the de Broglie relation:

$$mv = \frac{h}{\lambda};\ \text{thus } v = \frac{h}{m\lambda},$$

$$v = \frac{6.6 \times 10^{-34}\ \text{joule}\cdot\text{sec}}{10^{-30}\text{kg} \cdot 10^{-10}\ \text{m}} = 6.6 \times 10^{6}\ \text{m/sec}.$$

$$\left(\frac{\text{joule}\cdot\text{sec}}{\text{kg.m}} = \frac{\text{N}\cdot\cancel{\text{m}}\cdot\text{sec}}{\text{kg.}\cancel{\text{m}}} = \frac{\dfrac{\cancel{\text{kg}}\cdot\text{m}}{\text{sec}^{\cancel{2}}}\cdot\cancel{\text{sec}}}{\cancel{\text{kg}}} = \text{m/sec}\right)$$

20.12

The de Broglie wavelength is given by

$$\lambda = \frac{h}{mv} = \frac{6.6 \times 10^{-34}\ \text{joule}\cdot\text{sec}}{(0.2\ \text{kg})(1\ \text{m/sec})}$$

$$= 3.3 \times 10^{-33}\ \text{m}.$$

20.13

By definition, the de Broglie wavelength is given by $\lambda = h/mv$. But $KE = 1/2\ mv^2$, so

$$v = \sqrt{\frac{2(KE)}{m}}.$$

Hence the momentum $mv = m\sqrt{\frac{2(KE)}{m}} = \sqrt{2m(KE)}$;

substituting this for mv in the de Broglie relation, $\lambda = \dfrac{h}{\sqrt{2m(KE)}}$.

20.14

You would be unable to learn about the following:

a) objects outside your throwing range,
b) objects from which the ball would not bounce (sponge, etc.),
c) fluids like air,
d) objects appreciably smaller than the ball. Whereas you might learn of the existence of small objects, you couldn't learn the details of their shape. (Similarly, if you are to "see" detail with electromagnetic waves, the wavelength must not be much smaller than the dimensions of the detail.)

20.15

The uncertainty principle states

$$\Delta x \Delta p \geq \frac{h}{2\pi}$$

$$\geq \frac{6.6 \times 10^{-34}}{2\pi}\ \text{joule}\cdot\text{sec}$$

$$\geq 1 \times 10^{-34}\ \text{joule}\cdot\text{sec}.$$

The uncertainty in momentum is the product of the mass, 1×10^{-2} kg, and the uncertainty in speed, 3 cm/sec or 3×10^{-2} m/sec. That is,

$$\Delta p = 3 \times 10^{-4}\ \frac{\text{kg}\cdot\text{m}}{\text{sec}}.$$

$$\text{Thus, } \Delta x \geq \frac{1 \times 10^{-34}\ \text{joule}\cdot\text{sec}}{3 \times 10^{-4}\ \text{kg}\cdot\text{m/sec}}$$

$$\Delta x \geq 3 \times 10^{-31}\ \text{m}.$$

(This result indicates that uncertainty effects are completely negligible for objects of "normal" size and speed.)

20.16

The essence of the quantum theory is that energy exists in bundles of size proportional to the frequency. If the constant of proportionality were zero, the bundles would have no energy and there could be no quantum effect. Light would become entirely a wave phenomenon and there would be no photoelectric effect or compton effect.

20.17

Open-ended question, but the following comments are relevant.

Most physicists, probably all atomic physicists, regard with wonder (and, perhaps, also awe) the way in which Planck's constant appears in the description of atomic phenomena. This constant is one of the fundamental constants of physics, along with the charge on the electron and the speed of electromagnetic radiation in vacuum. We have seen it appear in the photoelectric effect and the Compton effect, in the energy hf and the momentum hf/c of radiation. It appears in the formulas that represent quantization of energy levels and of angular momentum. It also appears in the uncertainty principle, where it expresses the limitation on our ability to know both the position and momentum of a particle. An enormous number of atomic phenomena fall into a beautiful, consistent pattern that is held together rather remarkedly by Planck's constant.

20.18

a) $\lambda = 2(1\text{ m}) = 2\text{ m}$,

$$p = \frac{h}{\lambda} = \frac{6.6 \times 10^{-34}\text{ joule·sec}}{2\text{m}}$$

$$= 3.3 \times 10^{-34}\text{ kg·m/sec.}$$

But $p = mv$, so $v = p/m$,

$$v = \frac{3.3 \times 10^{-34}\text{ kg·m/sec}}{10^{-9}\text{ kg}}$$

$$= 3.3 \times 10^{-25}\text{ m/sec.}$$

b) Similarly, we can solve for v as above, or in one step as follows:

$$v = \frac{h}{\lambda m} = \frac{6.6 \times 10^{-34}\text{ joule·sec}}{2 \times 10^{-1}\text{ m} \times 6.6 \times 10^{-26}\text{ kg}}$$

$$= 5 \times 10^{-8}\text{ m/sec.}$$

c) $$v = \frac{6.6 \times 10^{-34}\text{ joule·sec}}{2 \times 10^{-6}\text{ m} \times 10^{-22}\text{ kg}}$$

d) $$v = \frac{6.6 \times 10^{-34}\text{ joule·sec}}{2 \times 10^{-10}\text{ m} \times 10^{-30}\text{ kg}}$$

$$= 3.3 \times 10^{6}\text{ m/sec.}$$

(In general, as the size of the particle and its containing "box" decreases, the least speed the particle could have increases!)

20.19

Open-ended question, but the following comments are relevant:

The uncertainty principle expresses limitations on our ability to obtain information about certain detailed and particular phenomena on the atomic scale. At this time, the application of the uncertainty principle to problems such as free will represents an extrapolation of physical theory that is unjustified in light of the absence of experimental facts.

The claims of these philosophers and physicists represents in part a reaction against the rigid determinism represented by Laplace's view expressed in S.G. 20.21.

The facts do not support a dogmatic stand on either side of this issue.

20.20

In one sense this statement is acceptable. If a scientific theory correctly predicts experimental results it is doing its job. In another sense, however, this statement falls short of those ends desired by most physicists. That is, in addition to a formalism (mathematical symbols and processes), a good theory has models in terms of which the physicist can "visualize" the implications of the theory. Such "visualizations" or "picturizations" can often lead to deeper understandings.

20.21

No. The precision with which the position and momentum can be determined is limited by the uncertainty principle. Thus, since the state of the particles of the universe is uncertain at any time, one cannot know the future with absolute certainty.

Of course, what a superior intelligence could or could not do is not within the scope of scientific discussion.

20.22

The idea of complementarity has been used by Bohr and other physicists to describe the use of the wave and particle pictures in atomic physics. The idea is another way of expressing human limitations in the study of atomic phenomena. Bohr has argued that complementarity must be used in other connections than atomic physics, but there has been small success, if any. So far the only uses have been in analogies, or in highly general statements. Some examples are:

Bohr's biological example: The vital aspect and the physiochemical aspect of living beings. The complete description, by means of physics and chemistry, of a living being would require an analysis so extreme in its various parts that it would inevitably lead to the death of the subject studied. The study of the vital functions would have to ignore in a large measure the details of the physical and chemical processes taking place in tissues and cells. Neither study by itself would give a complete description of the behavior of living beings.

Bohr's psychological example: We speak of living beings acting by instinct or with the use of reason. "Instinct" and "reason" seem to be mutually contradictory but complementary aspects of behavior.

An example of the use of the idea of complementarity in a more limited, perhaps trivial way: An account of what happens on a TV screen in terms of pictures of people and things is, in a sense, complementary to the description in terms of electron beams scanning the screen.

Prologue

One of the oldest, yet one of the exciting current problems of physics concerns the nature of matter. As long ago as the fifth century B.C. it was suggested that material things are made up of small, indivisible particles; yet in the early part of the twentieth century reputable physicists could still challenge the validity of the atomic theory of matter. This merger of a rich history with the ongoing of an incomplete story brings intrigue and excitement to the study of matter.

The Prologue to Unit 5 gives a brief sketch of the early theories of matter. All of these theories share in the search for an explanation of the multitude of microscopic changes in terms of a microscopic "stuff." The earliest thinkers argued whether the basic "stuff" was atomistic or continuous. Was there one basic "stuff" or were there several? Aristotle provided the answer which satisfied scholars for two thousand years.

The early theories of matter all interpreted macroscopic change as the consequence of some kinds of transformations at the microscopic (invisible) level. Thus several ingredients of modern atomic theory were present in these early theories of matter. Chief among these ingredients are (a) the forerunner of our concept of the element, and (b) the interpretation of change as a consequence of transformations among the "elements."

With these notions forming a vital part of man's view of matter, the stage was set for some very penetrating questions. Many of the techniques developed by the alchemists were available as a means for carrying out experiments. These ideas, these questions and these techniques, in the hands of Boyle, Lavoisier and Dalton, were the beginning of modern chemistry.

Sec. 17.1 Dalton's atomic theory and the laws of chemical combination

The historical background for this section is given in the Prologue to Unit 5. It should be pointed out that the work of the eighteenth century chemists provided important evidence for the atomic theory. Two accomplishments of these scientists were essential to the work of Dalton. First was the establishment of the concept of element. This concept, which has its roots in antiquity, was sharpened by Boyle in 1661 and was brought close to its modern form by Lavoisier in 1789. Second was the establishment of quantitative methods as the approach to chemical problems. This latter point was particularly significant.

The postulates of Dalton's atomic theory reflect his confidence in the validity of the law of conservation of mass. As Dalton wrote,

> No new creation or destruction of matter is within the reach of chemical agency. We might as well attempt to introduce a new planet into the solar system, or to annihilate one already in existence, as to create or destroy a particle of hydrogen.

Perhaps the most important contribution of Dalton was his emphasis on the weights of atoms. Again to quote Dalton:

> In all chemical investigations it has justly been considered an important object to ascertain the relative weights of the simples which constitute a compound.

(Simples as used by Dalton is equivalent to elements.)

Sec. 17.2 The atomic masses of the elements

The work of the early chemists can be illustrated by the following analogy. Suppose you are a guest at a tea party. Your hostess challenges you to determine whether granulated sugar or cube sugar is being used by the guests in the adjoining room. You accept the challenge and request your hostess to select five cups of freshly prepared tea from her guests. After evaporating the liquid you weigh the sugar residue. You find the weight to be a multiple of some unit, say 1.4 units. You conclude, of course, that cube sugar was used. Note that your conclusion is drawn without having ever seen a sugar cube!

The conclusions about relative atomic masses are drawn in a similar, but more indirect manner. The indirectness follows from the fact that several types of experimental results must be utilized to arrive at the relative atomic masses.

If your students desire a more detailed description of the way in which relative atomic masses were determined, wait until Sec. 17.8 is concluded. Three lines of evidence can then be utilized. Note: Pursuing the evidence upon which relative atomic masses were based is not recommended; therefore, the material in the appendix of this chapter is presented only to enable you to give guidelines to interested students. The PSSC film number 0110, Definite and Multiple Proportions, would be useful in providing background for this section. L.K. Nash has written a case study of the atomic-molecular theory which is published in the Harvard Case Histories in Experimental Science. A paperback, Through Alchemy to Chemistry, by J. Read, a Harper Torchbook, gives historical background for Chapter 17.

Sec. 17.3 Other properties of the elements: valence

Dalton's theory provided an explanation for the conservation of mass and the law of definite proportions. Yet many questions were left unanswered. Chief among these was the question, "What makes atoms unite with each other?" The idea of affinity was introduced in an attempt to "explain" why nitrogen will react with more hydrogen than will titanium. Unfortunately, the idea of affinity explained nothing. It is equivalent to saying that wood burns because it is combustible.

The concept of valency slowly developed as an attempt to understand and therefore to predict the way in which atoms would combine. Valence numbers could be assigned by an analysis of data such as occurs in Tables 17.2 and 17.3. With these valence numbers one could predict the outcome of a reaction between two elements; however, one still could not understand why atoms united. Understanding did not come until the atom took on an internal structure.

A note of warning! The word valence came to mean different things to different chemists; consequently, it has been abandoned in some modern chemistry texts. However, even where the noun form is rejected, the adjective form still finds broad usage. The bonding electrons are called valence electrons and these electrons are found in valence orbitals. A frequently used near-synonym of valence is oxidation number.

Sec. 17.4 The search for order and regularity among the elements

Two developments stimulated the discovery of new elements. The first was the precise definition of an element framed by Lavoisier. This definition suggested new experiments, for example, the investigation of gases, which earlier had all been considered to be the same element. The second was the development of new physical techniques which aided in the discovery of new elements. Electrolysis (Sec. 17.7 and 17.8) and spectroscopy (Sec. 19.1) were experimental techniques which assisted in the quest for new elements. The periodic table, an empirical relationship found to exist among the elements, suggested the properties of unknown elements and hence hastened their discovery (Sec. 17.5).

Among the elements, there were found those which possessed very similar properties. Classifications were made. Here is an example of the fact that classification is a very important kind of scientific activity. It represents the first step in bringing order out of chaos. Frequently the development of a theory is preceded by the classification of facts. Two such theories, to be considered later, deal with atomic spectra and radioactive series. A number of broad classifications can be made by even the casual observer: organic vs. inorganic, metal vs. non-metal, solid vs. liquid vs. gaseous, dense vs. porous, etc. These relationships found to exist among the elements set the stage for the next development.

Sec. 17.5 Mendeleev's periodic table of the elements

The accumulation of data concerning the properties of the elements and the discovery of relationships between the elements gave rise to a more general relationship based on the atomic weights of the elements. In Mendeleev's own words,

> The law of periodicity was a direct outcome of the stock of generalizations and established facts which had accumulated by the end of the decade 1860-1870; it is an embodiment of those data in a more or less systematic expression.

The periodic table is the basis for the periodic law which states that when the elements are arranged (ordered) according to their atomic masses, a periodicity of their properties results. At about the same time that Mendeleev published his version of the periodic table, Meyer, in Germany, enunciated the periodic law while in England Newlands noted the repetition of properties when the elements are arranged in order of atomic weights. Why then is Mendeleev credited with the discovery?

The reason is that Mendeleev did much more than produce an arrangement of the elements. Generally, the periodic law is regarded as growing out of the periodic table; however, as far as Mendeleev was concerned, perhaps the opposite was true. It appears that he was convinced of the validity of the periodic law and his arrangement of the periodic table merely conforms to it. Thus, titanium had to come in the same family as silicon. This represents a departure from the schemes proposed by his contemporaries.

Furthermore, the placement of titanium under silicon left a vacancy under aluminium. This provided Mendeleev with an opportunity—the opportunity to use the periodic law to deduce the properties of an unknown element. This clinched it! Given the alternatives, the scientific community will always choose a theory that not only correlates data, but also predicts new results. Again to quote Mendeleev:

> The confirmation of a law is only possible by deducing consequences from it, such as could not possibly be foreseen without it, and by verifying these consequences by experiment.

The book The Discovery of the Elements by Mary Elvira Weeks would be useful for Sections 17.5 and 17.6. This book, published by the Journal of Chemical Education, contains an extensive bibliography.

Sec. 17.6 The modern periodic table

In 1894, Sir William Ramsay wrote to Lord Rayleigh: "Has it occurred to you," he wrote, "that there is room for gaseous elements at the end of the first column of the periodic table?" Prior to this letter, Ramsay had been working with atmospheric nitrogen. He found that a small fraction of residual gas was left after the nitrogen had been absorbed by hot magnesium. A spectroscopic analysis of the gas showed it to be a hitherto unknown constituent of air. The gas was argon. Thus began a series of discoveries which resulted in a major modification of the Mendeleevian version of the periodic table.

Helium had been "known" since 1868. In 1868 helium was observed spectroscopically as a constituent in the sun's chromosphere. It went unexplained until Ramsay, in 1894, extracted small amounts of gas from uranium ore and discovered terrestrial helium. With helium (atomic weight 4.0) and argon (atomic weight 40) discovered, the periodic table is again left with a vacancy. In 1897, Ramsay, speaking in Toronto, stated:

> There should therefore be an undiscovered element between helium and argon with an atomic weight 16 units higher than that of helium and 20 units lower than that of argon... and pushing this analogy further still, it is to be expected that this element should be as indifferent to union with other elements as the two allied elements.

Thus, neon was predicted!

The modern periodic table is one of the most useful devices to the chemist. In experienced hands, it can lead to new discoveries of many kinds. Haber used the table in the development of his high pressure catalytic synthesis of ammonia, an important industrial process. A more recent example is the discovery of the freons which are important gases for use in refrigerators and air-conditioning units. In a matter of a few hours after determining the desirable qualities of these refrigerants, Thomas Midgley Jr. and two associates deduced from the relationships of the periodic table that the substance CCl_2F_2, a freon, should have the desired properties of being stable, nontoxic, nonflammable, noncorrosive, etc. This discovery led to the development of a large industry.

By 1900, a fundamental basis for the periodic law was needed. At this time, the periodic table was an empirical device. As an empirical device it was very useful—just as Kepler's laws were useful for calculation purposes. Until the inverse-square law was postulated by Newton, however, Kepler's laws remained a mystery. Likewise, until the internal structure of the atom was studied, the periodic table remained a mystery.

Sec. 17.7 Electricity and matter: qualitative studies

Here we see another great synthesis: two subjects, previously thought to be unrelated, are beginning to coalesce. Some of the other syntheses that have occurred are terrestrial physics and celestial physics, electricity and magnetism, electromagnetism and light, heat and matter, and space and time (Chapter 20). Syntheses always lead to a deeper understanding of phenomena. Here the synthesis is between electricity and matter — a link is established between them.

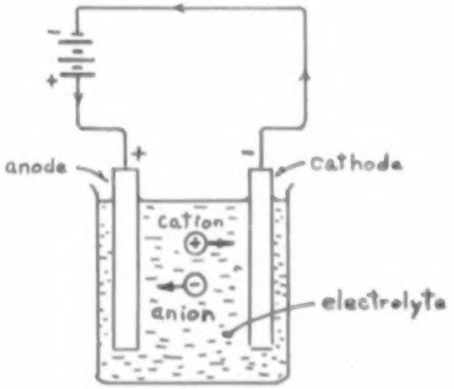
All substances seem to fall into one of two categories: those that allow an electric current to pass through them with ease and those that do not. The former we call conductors and the latter insulators. Conductors can be gaseous, liquid, or solid. Ordinarily, gases are poor conductors; however, when they are subjected to a high potential or to certain kinds of radiation, they become highly conductive. The solid conductors are the metals. (An important class of materials such as silicon and germanium also conduct electricity, but with difficulty. Such materials are called semi-conductors.) Liquids that are electrical conductors are called electrolytes.

Among the first investigators of the interrelationships between electricity and matter were Humphrey Davy and J.J. Berzelius. Growing out of their work was one of the first attempts to systemize chemical behavior. Berzelius developed the dualistic theory which was to form the basis for chemical combination. He assumed that chemical and electrical attraction were essentially the same. Atoms were believed to be polar. Chemical combination was the result of the interaction of the polar atoms. This theory, advanced in 1812, was important until the advent of organic chemistry in the 1830's. Then the dualistic theory became a handicap rather than a help.

Sec. 17.8 Electricity and matter: quantitative studies

In the last section it was established that chemical changes are brought about by electricity. This was conclusively demonstrated by Davy. Later Faraday, who started his scientific career as an assistant to Davy, established quantitatively the amount of chemical change caused by a given quantity of electricity.

One of Faraday's important contributions was the development of the descriptive terms which are now universally used. He, together with William Whewell, devised the terminology in 1833 which enables one to describe the mechanism of electrolysis. The conductor, solution, or molten salt, is the electrolyte.* The conductors by which the positive current enters or leaves the electrolyte are the electrodes; the positive current enters the positive electrode, the anode, and leaves the negative electrode, the cathode. The charged particles that move toward the anode are called the anions and those that move toward the cathode are called cations.

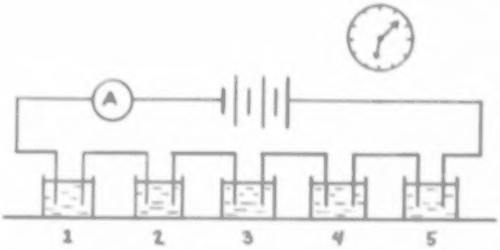


The Process of Electrolysis

In the process of electrolysis, interesting energetics are involved. The battery converts chemical energy into electrical energy; that is, a potential difference is established between the terminals making the anode positive with respect to the cathode. Thus, positive particles are attracted to the cathode and negative particles to the anode. Where does the electrical energy go? It is dissipated as heat in the electrolyte and in the external circuitry.

Quite early in his research, Faraday believed that a quantitative relationship could be established between chemical change and quantity of electricity. His efforts to discover this relationship were frustrated by the fact that secondary reactions often occurred to complicate the results. However, after he developed the volta-electrometer, since 1902 called the coulometer, his work proceeded more smoothly. The coulometer served as a standard. When it was placed in series with other electrolytic cells, the amount of decomposition occurring in them could be compared to the amount of hydrogen liberated.

A typical experimental arrangement for verifying Faraday's laws of electrolysis is shown below. Five beakers are placed in series with a battery and an ammeter. A timing device is also necessary. If the experiment is conducted in such a way that the product of current and time equals 96,540 coulombs, data such as shown in Table 17.1 will result. The relative masses have the same ratios as those determined from chemical analysis. Thus, there is a quantitative connection between chemical change and amount of electricity.



Beaker 1: a coulometer
Beaker 2: HCl solution
Beaker 3: molten NaCl
Beaker 4: molten $MgCl_2$
Beaker 5: molten $AlCl_3$

This kind of experiment greatly assisted the chemists with one of the big problems of their day, atomic weight values. Under explicit conditions, the amount of an element liberated at an electrode is called the equivalent mass. In Faraday's own words:

> I have proposed to call the numbers representing the proportions in which they are evolved electro-chemical equivalents.

Thus, in the table above, the equivalent masses of H, Na, Mg and Al are 1, 23, 12 and 9. The equivalent masses of these elements determine how much of them will combine with other elements.

There is a profound implication of

---

*All underlined words were devised by Faraday and Whewell.

TABLE 17.1

	Anode		Cathode			
Beaker	Element	Wt. released (gram)	Element	Wt. released	Atomic Weight	Total Charge Passed (coulombs)
1	O_2	8.00	H	1.008	1.008	96,540
2	Cl	35.5	H	1.008	1.008	96,540
3	Cl	35.5	Na	23.0	23.0	96,540
4	Cl	35.5	Mg	12.16	24.32	96,540
5	Cl	35.5	Al	8.99	26.98	96,540

the laws of electrolysis. This implication was eloquently expressed by Helmholtz in 1881 in his Faraday Lecture ture delivered at the Royal Institution.

> Now the most startling result of Faraday's law is perhaps this: if we accept the hypothesis that the elementary substances are composed of atoms, we cannot avoid concluding that electricity also, positive as well as negative, is divided into elementary portions which behave like atoms of electricity.

As we shall see in the next chapter, it was just 16 years after "atoms of electricity" were proposed by Helmholtz that the electron was discovered ered.

Thus we have gone a full circle. We have used electricity to gain insight into the nature of matter and we have succeeded. However, our success was greater than we anticipated, for in the process we gained insight into the nature of electricity. The stage is all set for the next part of our story.

Sec. 18.1 The problem of atomic structure: pieces of atoms

Historically, one might say that physics began with Galileo and chemistry with Lavoisier and Dalton. It is important to stress at this point the intermingling of chemistry and physics. The concept of element postulated by Lavoisier became inseparable from Prout's law (discussed in Chapter 23). Prout's hypothesis (1815) of basic building-up units of matter found acceptance in latter-day physics and chemistry. Mendeleev's periodic table revealed a need for basic "structure" on which atoms might be built. However, a lack of any direct experimental evidence pointing toward a structure for atoms prevented further speculation. No reasons could be advanced at the time for the periodicity of elements with similar properties. It was left for later experimenters like Becquerel and others finally to find the way, and not until the theory of atomic structure was advanced by Bohr did the periodic table find complete acceptance and experimental support in the microscopic domain. Suggested Reading: From Atomos to Atom, a Harper Torchbook by Andrew G. Van Melsen.

Sec. 18.2 Cathode rays

The pioneer work of Geissler and Plucker led to subsequent discoveries connected with cathode rays. However, one must not forget to give due credit to Crookes. Crookes' new interest in vacuums led him to study the Geissler tubes. He perfected them for more efficient study of the radiation, and they have been called "Crookes' tubes" ever since.

Crookes represented the results dramatically and systematically. He showed that cathode rays travel in straight lines and can cast shadows. He also showed that the radiation can turn a small wheel when it strikes one side. Crookes also showed that the radiation can be deflected by a magnet. He was convinced, therefore, that he was dealing with charged particles and not electromagnetic radiation. Crookes spoke of these charged particles as a fourth state of matter, or an ultra-gas, as far beyond the ordinary gas in rarefaction and intangibility as an ordinary gas is beyond a liquid.

Crookes on several occasions nearly stumbled onto great discoveries that were eventually made by others. (More than once he fogged photographic plates during the running of his tube, though his plates were contained in their boxes. However, he missed the connection and it was Roentgen who later, using Crookes' tube, discovered the x rays.)

The Aurora Borealis phenomenon has been the subject of speculation ever since the time of Benjamin Franklin, who attributed the phenomenon to electricity. That it is due to electric rays from the sun was suggested in 1872 by A.B. Donati of Florence. Eugen Goldstein of Berlin held that they were cathode rays from the sun. Kristian Birkeland (1867-1917) of Christiania (now Oslo), adopting this view, constructed a miniature model of the earth (terrela) and exposed it to cathode rays in a vacuum tube. Later when this terrela was magnetized, it possessed an illumination concentrated upon a spiral path about the poles and a thin luminous ring about the equator. The mathematical theory of these phenomena was elaborated by Carl Stormer.

During the last decade of the nineteenth century the scientific world was divided over the nature of cathode rays. The English school favored Crookes' theory that cathode rays consisted of tiny negatively charged particles. The German investigators were unanimously opposed. The German school was of the opinion that cathode rays were ether waves similar to the electromagnetic (radio) waves discovered by Hertz in 1887. This view was strengthened by Hertz's discovery that cathode rays were small enough to penetrate gold leaf.

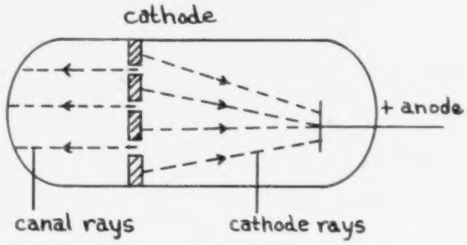
If cathode rays really consisted of negatively charged particles, they would be deflected both by an electric field (electrical force) and by a magnetic field. Hertz was unable to detect such an effect in an electric field, no matter how he tried to perform this experiment. Thomson decided to repeat Hertz's experiment and also to remove what he thought was a discrepancy in earlier procedures.

When Thomson tried to deflect cathode rays by passing them between electrically charged plates within the vacuum tube, he also obtained no result. Although he observed a slight flicker when the electric field was first turned on, he discovered that he could not obtain a permanent deflection no matter how strong he made his electric field.

After much thought Thomson finally theorized that the cathode rays were converting the particles of gas into charged ions upon collision, and these charged ions were then immediately attracted to the plate of opposite charge. Thus the plates were neutralized by the ionized gas particles almost instantaneously and they could no longer produce an electric field.

The remedy was to have the gas in the discharge tube at a very low pressure. When this was done, Thomson was able to detect his beam of cathode rays being bent in the electric field. He obtained rays with the fantastically high velocity of 160,000 miles per second in subsequent experiments. His conclusions were then incontestable.

Cathode rays were first brought to notice in 1855. It may be helpful to mention here that in 1886 Goldstein discovered a new kind of radiation in the Crookes' tube which he called "canal rays." These are discussed in some detail in Unit 6. If the cathode is made of a perforated metal plate (i.e., one with holes in it), then, in addition to the cathode rays traveling from the cathode to the anode, we also observe a stream of rays behind the cathode. When the charges of these rays were found by deflection in electric and magnetic fields, they were discovered to be positive and integral multiples of elementary charges. If q_e = the elementary charge, then only integral multiples of q_e appear, i.e., nq_e, where $n = 1, 2, 3$...or some whole number.



The masses of these positive rays depend on the nature of the substances in the tube. When hydrogen is used in the tube, the multiple of integral charge is always $n = 1$. This was, in fact, one of the first instances of the direct evidence of singly charged positive ions of hydrogen. In a sense, it was an "alter ego" of the electron.

The positive rays appear only when there is enough gas in the tube so that ionization can be produced by the collisions due to the cathode rays.

It may be significant at this point to mention positive rays so that later concepts of a positive center and negative orbit for a neutral atom in the Bohr theory may find easy acceptance.

Sec. 18.3 The measurement of the charge of the electron: Millikan's experiment

Among the several successful attempts made to measure the charge on an electron, prominent are those of J.S. Townsend, J.J. Thomson, H.A. Wilson and R.A. Millikan, several of them familiar today as Nobel Prize winners. Although Millikan's method was probably the most accurate and the simplest, he was not the only one who carried out experiments on the charged particles. From his measurements he demonstrated what had been previously surmised by Benjamin Franklin and by more recent physicists, namely, that electricity has a corpuscular structure.

As early as 1899 it was shown by John S. Townsend of Oxford that the positive or negative charge carried by an ion in a gas was equal to the charge carried by the hydrogen ion in the electrolysis of water. Millikan showed conclusively that electricity consists of equal units, that the electric charge of each single ion is always a multiple of this unit, and that this unit of charge is not merely a statistical mean, as the atomic weights have been shown to be.

For reasons of historical interest let us consider Thomson's measurement of the electronic charge. He had already performed experiments yielding q_e/m values for the cathode rays, as explained in the text. Though the results were consistent enough, Thomson was by no means certain that the difference between the q_e/m value for cathode rays and the q_e/m value for hydrogen ions in electrolysis was entirely due to the enormous difference in the relative masses of the hydrogen ion and the "corpuscle"—i.e., electron. Indeed, the charge on the negative electron could be several times that of the hydrogen ion and still leave the latter very much heavier than the former.

The only way to resolve this was to determine q independently. Thomson's method was described in a paper published in the Philosophical Magazine, December 1898. He used ions in a gas which acted like nuclei for condensation from supersaturated water vapor, thus producing a cloud of water droplets. One can determine the charge carried by each of them, and hence by the original ions in the following way.

By measuring the downward velocity of the cloud falling under gravity, he calculated the radius of each droplet. This was done by using Stoke's formula for viscosity:

$$V = \frac{2(\rho - \sigma)ga^2}{9\mu}$$

μ = coefficient of viscosity

ρ = the density of drops

σ = density of the air

g = gravity

a = radius

V = velocity of the drop

Next the cloud was made to move under the influence of an electric field and the ions fell on a plate connected to a condenser. The rate at which the potential on the condenser changed was measured. Since the number of ions originally present could be calculated by knowing the amount of water which was condensed on all the ions, the charge on a single ion could be calculated. The results showed large variations, and the method was obviously inapplicable to the problem of determining the charge carried by individual corpuscles.

Contrast this with the simplicity and accuracy of Millikan's method. The smallest charge that Millikan measured had the value 1.6×10^{-19} coulombs. All other electric charges were multiples of this charge. More recently discovered positrons have a charge equal in magnitude to that of electrons but opposite in sign. Since then particles of various masses have been discovered, but so far all have charges that are integral multiples of the electronic charge. Likewise, recently discovered particles called mesons of masses 273 times that of the electron also have just one electronic charge.

There is no direct correlation between the mass of a particle and its charge. The electronic charge has been found on some of the lightest as well as some of the heaviest particles in existence, for example, a positron and a proton.

Millikan's method has now been outdated by visual methods using bubble chambers, etc. where the tracks of particles can be actually observed as bent in magnetic fields and direct inferences can be drawn.

Sec. 18.4 The photoelectric effect

The photoelectric effect was discovered by Hertz in the course of his work designed to show experimentally that Maxwell's prediction of electromagnetic waves was correct. In particular, the wave theory of light had been incorporated in Maxwell's electromagnetic theory, and thus all observable phenomena were explained in terms of it. The example indicated below serves to illustrate why the wave theory in its electromagnetic form was incapable of giving the correct order of magnitude for the time taken in ejecting electrons.

Derivation of time delay required by wave model

In a properly lit classroom there are at least 10 foot-candles of illumination at the laboratory tables. This is equivalent to about 1500 ergs/sec/cm^2 of radiant energy.

Suppose that these 1500 ergs/sec fall on 1 cm^2 of zinc causing it to emit photoelectrons.

In zinc, the atoms are 2.5×10^{-8} cm apart. Thus in the top layer of zinc atoms in 1 cm^2 there are:

$$\frac{1}{(2.5 \times 10^{-8})^2}$$

$$= 1.6 \times 10^{15} \text{ atoms/cm}^2.$$

If the light consists of waves that penetrate ten layers of atoms, its energy will be distributed over $10 \times 1.6 \times 10^{15} = 1.6 \times 10^{16}$ atoms. Now suppose that each atom absorbs an equal share of the 1500 ergs/sec which are available. (Notice how our wave model differs from our particle model interpretation on this point.)

This is:

$$\frac{1.5 \times 10^3 \text{ ergs/sec}}{1.6 \times 10^{16} \text{ atoms}}$$

$$= 9.4 \times 10^{-14} \text{ ergs/sec/atom.}$$

The work function of zinc is about 4.9×10^{-12} ergs.

To acquire this much energy at the rate of 9.4×10^{-14} ergs/sec an atom of zinc would have to "save up" for

$$\frac{4.9 \times 10^{-12} \text{ ergs}}{9.4 \times 10^{-14} \text{ ergs/sec}} = 52 \text{ sec}$$

before it had enough energy to emit a photoelectron.

The work functions of some other materials yield times twice as long under similar conditions.

Sec. 18.5 Einstein's theory of the photoelectric effect: quanta

Einstein's theory of the photoelectric process can be suitably explained in terms of the illustrative analogy shown in Fig. 18.1. The illustration explains clearly the notion of work function and threshold frequency. A quantum "kick" of energy hf sends the electron up a height W and provides some additional kinetic energy as well. This explains the character of W as potential energy.

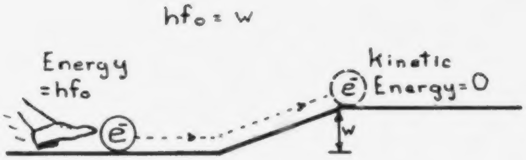


Fig. TG 18.1

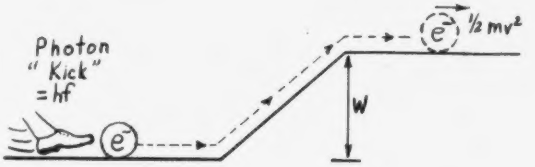


Fig. TG 18.2

In Fig. 18.2 the idea of threshold energy is explained. The kick is just sufficient to overcome the potential height W, so that the electron at the top of the hill has zero kinetic energy. It is obvious that any energy less than hf_o, (which is just enough to send the electron up the hill), will fail to eject the electron over the potential hill.

Sec. 18.6 X rays

X rays are similar to gamma rays; however, the two have a different origin. X rays derive from atomic electrons whereas gamma rays originate in the atomic nucleus. Gamma rays have discrete energies whereas x rays from a conventional x-ray tube have a continuous energy spectrum over a considerable range of wavelengths. The "peak voltage" is used to characterize both the voltage at which the tube is operated and the maximum energy of the resulting x rays.

Supplementary X-ray Information

Frequency cycles per sec	Wavelength cm.
10^{16}	3×10^{-6}
to	to
3×10^{20}	10^{-10}

The frequencies of gamma rays and x rays overlap at about 10^{18} sec^{-1}, where the lower limit of gamma-ray frequencies lies. Radiation of wavelength 3×10^{-6} cm could be produced by a voltage of 100 volts across an x-ray tube, and would be called "soft" x rays. Wavelengths of the order of 3×10^{-9} cm are obtainable from tubes using 100,000 volts.

Identical radiation, called gamma rays. is also obtainable from nuclei. When electrons are decelerated, thus producing very high frequency (or very small wavelength) radiation, it is difficult to decide whether to characterize the radiation as x-ray or gamma-ray. Electrons with energies up to a billion electron-volts can produce such high-frequency radiation. Actually the term x rays is reserved for lower frequency radiation produced by the x-ray tube.

The decision to discuss gamma rays at this point is left to the discretion and inquiry of the teacher and student respectively. Since they are very similar types of radiation, there may be an advantage in this type of approach insofar as a brief comparison is concerned.

Brief Note on Roentgen

Wilhelm Konrad Roentgen received the Nobel Prize in 1901. It is said that he objected strongly to calling his x rays by the name "Roentgen rays." He believed that scientific discoveries belonged to mankind and that they should not in any way be hampered by patents. In fact, he chose to donate his Nobel Prize to the University of Wurzburg. Roentgen died on February 10, 1923 in his seventy-eighth year. Ironically, he died of cancer, a disease which often responds to treatment by x rays.

Sec. 18.7 Electrons, quanta and the atom

Directly after the discovery of cathode rays (electrons) and the indirect evidence for positively charged matter, speculations were made about the nature of neutral atoms. A number of theories of atomic structure were proposed in the period 1897 to 1907. The most prominent of all these theories was that advanced by J.J. Thomson. His theory made an attempt to meet the following requirements:

(a) The atom had to be a stable configuration of positive and negative charges.

(b) The atomic theory had to offer some explanation for the details of atomic spectra.

(c) The theory had to account for the chemical differences and resemblances of elements.

Thomson's interest in the relation between chemical properties and atomic structure may well have derived, in part at least, from the indifferent attitude of chemists to his discovery of the electron.

Thomson's model had to deal with a collection of positive and negative charges in equal number, where the charges obeyed the inverse-square law of attraction and repulsion. These charges would be required to settle in a position of stable equilibrium, without falling into one another. Oscillations of these charges about this stable position would result in the emission of line spectra. The problem was not an easy one, and was eventually explained differently (See Sec. 19.5.)

J.J. Thomson proposed his theory in 1904. His model had stable distributions of rings of different numbers of electrons rotating within a sphere of positive electricity. He made an attempt to correlate his stable configurations to chemical properties as seen in Mendeleev's table. It may be worth mentioning at this point that this picture of the atom was not due to Thomson alone, and was, in effect, based on an idea previously suggested by Kelvin. This was published in the Philosophical Magazine, Vol. 3, p. 257, 1902, entitled "Aepinus Atomised" in which Kelvin developed a theory of electricity based on the properties of electrons. J.J. Thomson's paper was first seen in 1904 in The Philosophical Magazine, Vol. 7, p. 237, entitled: "On the structure of the atom: an Investigation of the stability and Periods of Oscillation of a number of corpuscles arranged at equal intervals around the circumference of a Circle; with Application to the results of the Theory of Atomic Structure."

In his mathematical treatment he considered corpuscles (electrons) at rest within the positive sphere, and also corpuscles in angular motion about the center of the sphere. He came out with various configurations for various angular velocities and states of rest. The fundamental idea of his theory was that the atom consists of a number of corpuscles moving about in a sphere of uniform positive electricity. This theory raises at least three questions:

(1) How do the corpuscles arrange themselves in the sphere?

(2) What properties does this structure confer on the atom?

(3) How can this model explain the theory of atomic structure?

The details of Thomson's model are mathematical and much too complicated for class discussion. The model survived at the time because it seemed to offer more possibilities for further development than did other theories. A quotation from Thomson's original paper is as follows:

> The analytical and geometrical difficulties of the problem of the distribution of the corpuscles when they are arranged in shells are much greater than when they are arranged in rings, and I have not, as yet, succeeded in getting a general solution.

Sec. 19.1 Spectra of gases

During the latter part of the nineteenth and into the early years of the twentieth century, two lines of interest were converging. The first was the interest in the nature of matter. With the discovery of the electron and radioactivity, new questions about the internal structure of the atom were being asked. The Thomson model of the atom was one attempt to answer some of these questions.

A second line of interest was the study of spectra. By the beginning of this century, a vast amount of spectroscopic data had accumulated. It was known that each element possessed a unique spectrum. Certainly, any proposed theory of atomic structure had to account for the origin and characteristics of spectra.

Since no experimental technique has contributed more to our understanding of the intrinsic structure of atoms and molecules than spectral analysis, a brief description of the technique will be given in the article section, page 90.

Sec. 19.2 Regularities in the hydrogen spectrum

In the Balmer formula we once again encounter an empirical relationship. It is important to see the role played by such relationships. First, an empirical formula should never be identified with an explanation, or theory. Second, while empirical formulas do not provide understanding, they simplify and clarify what the theory must explain. Newton's inverse-square law very quickly took on significance when the known laws of Kepler could be deduced from it. Likewise, Bohr's model of the atom was enhanced when Balmer's formula could be deduced from it.

Sec. 19.3 Rutherford's nuclear model of the atom

During the year 1908, Rutherford and his associates, Geiger and Marsden, initiated experiments on the scattering of alpha particles by a thin metallic foil. In 1909, they observed to their surprise that alpha particles could be scattered through a large angle (>90°). The scattering experiments were completed in 1909; however, Rutherford pondered on their significance for a long time. Early in 1911, Geiger relates that

> One day Rutherford, obviously in the best of spirits, came into my room and told me that he now knew what the atom looked like and how to explain the large deflections of alpha particles. On the very same day I began an experiment to test the relations expected by Rutherford between the number of scattered particles and the angle of scattering.

In order to understand the drama associated with this discovery, it would be effective to recall the then current thinking on atomic structure. In the Thomson model of the atom, the mass of the atom was distributed uniformly through the volume of a sphere. With such an atom only small angle deflections of the alpha particle should be observed. An analogy can serve to show this quite convincingly. Let us think of the Thomson atom as a marshmallow. A bullet (our "alpha particles") incident upon layers of such "atoms" would suffer little deflection. Since the Thomson model of the atom was the most popular with physicists, they were indeed surprised when alpha particles were observed coming backwards.

When Rutherford informed Geiger that he knew what the atom looked like and could explain the scattering results, what was the basis for experimental verification? In the first place, Rutherford envisioned a new model of the atom—the nuclear model: a massive, positively charged nucleus surrounded by planetary electrons. (Actually an earlier nuclear model had been proposed by H. Nagaoka, a Japanese physicist. Nagaoka's model derived its inspiration from the planet Saturn. He envisioned electrons traveling in rings about a massive center forming a miniature Saturn-like system.) In its totality, however, the atom is mostly empty space. Now instead of the gold foil being thought of as layers of marshmallow-like atoms, it becomes an array of widely spaced massive nuclei.

Rutherford scattering is further discussed in the article section, page

A very interesting biography of Rutherford has been written by E. N. da C. Andrade. It is entitled Rutherford and the Nature of the Atom and appears in the Science Study Series.

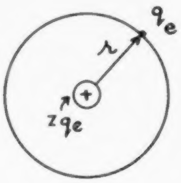
Sec. 19.4 Nuclear charge and size

Rutherford's scattering experiments were the beginning of an experimental technique that has been one of the most fruitful in producing information about the nucleus. Even today, scattering experiments are a widely used technique. By 1912, a rather precise model of the atom had emerged. The scattering experiments had indicated that:

1) the mass of the atom was concentrated in a very small volume relative to the atomic volume. This was indicated by the fact that most alpha particles suffered no deflection whatsoever in passing through the foil.

2) the concentration of mass carried a positive charge. This was suggested by the fact that of those alpha particles scattered, a very large fraction were scattered through small angles.

3) the size of the nucleus was on the order of 10^{-14} meter in diameter. This can be deduced by calculating the distance of closest approach of the alpha particle to the nucleus. The calculation is made by assuming that all the initial kinetic energy of the alpha particle is converted into electrostatic potential energy at the distance of closest approach. Thus, knowing the masses of the alpha particle and the scattering nucleus and knowing the initial kinetic energy of the alpha particle, the nuclear diameter can be deduced.



The magnitude of the nuclear charge was found to be equal to the product of the atomic number and the electronic charge. Thus, the arrangement of the elements in the periodic table, arranged in terms of atomic number, could now be related to atomic structure.

Moseley, who was killed during World War I, provided convincing evidence for the importance of the atomic number. He systematically followed up a discovery made by Bragg in 1913 that the heavier elements, when strongly excited, exhibit characteristic lines lying in the x-ray region of the spectrum. These x-ray spectra are quite simple as the figure below shows. Each element gave peaks having a slightly different wavelength. There was such regularity that Moseley was able to express the results in the form of an empirical formula similar to the empirical formula of Balmer. This formula may be written

$$f = 2.48 \times 10^{15} (Z - 1)^2$$

where f is the frequency in cycles per second and Z is the atomic number. This equation led to the discovery of new elements. For example, when the known elements were arranged according to increasing frequency of their x-ray lines, a gap existed at $Z = 43$, indicating the existence of an element (now called technetium) then unknown.

In spite of its usefulness, the nuclear model was not without its difficulties. As a review of some of the mechanics learned in Unit I, let us analyze the motion of a planetary electron. The centripetal force maintaining the electron in its circular orbit about the nucleus is supplied by the electrostatic attraction between electron and nucleus. Thus we can write

electrostatic force = centripetal force

or

$$9 \times 10^9 \frac{Z^2 q_e^2}{r^2} = \frac{mv^2}{r} .$$

Classically, an accelerated charge produces electromagnetic waves. Since an electron moving in a circular orbit is constantly accelerating, it should radiate. The expected frequency of the electromagnetic waves is just the frequency of the electron's motion about the nucleus. What is the frequency of the electron's motion?

frequency = number revolutions/second

$$f = v/2\pi r.$$

From our relation above

$$v/r = (9 \times 10^9 \; Zq_e^2/mr^3)^{\frac{1}{2}}$$

so

$$f = (1/2\pi)(9 \times 10^9 \; Zq_e^2/mr^3)^{\frac{1}{2}}.$$

If one assumes $Z = 1$ and $r = 0.5$ Å, one finds that $f = 7 \times 10^{15}\ \text{sec}^{-1}$.

Thus the atom should be emitting ultraviolet radiation! If it does, it loses energy. If it loses energy, r gets smaller, the electron makes more trips around the nucleus per second making the frequency of emitted radiation even higher. On this basis the atom should collapse in less than 10^{-8} seconds! The question is, how do we account for the stability of the atom?

Sec. 19.5 The Bohr theory: the postulates

Before discussing the Bohr theory of the atom, it might be wise to list the questions to be explained by any atomic model.

1. valence: what determines the ability of an atom to combine with other atoms?

2. periodic law: what is at the base of the family relationships?

3. periodic table: can any model give insight to the ordering of elements as they are in the periodic table?

4. electrolysis laws: would an under-

standing of valence provide an understanding of the laws of electrolysis?

5. scattering data: how can a nuclear model be stable?

6. spectra of elements: what is the origin of spectra?

All of the consequences of the Bohr theory can be logically deduced from his basic postulates. An alternative way of stating his postulates follows:

Postulate 1. The electron can exist only in certain stable, circular orbits in which the electron obeys the laws of mechanics. (Here the word stable means that the electron does not lose energy by radiating.) When the electron is in a stable orbit, the atom is said to be in a stationary (i.e. stable) state.

Note: This postulate avoids the difficulty discussed in Section 19.4 that the atom should be unstable. With this postulate, the atom is stable by definition!

Postulate 2. An atom can undergo a transition from one stationary state to another stationary state and in so doing emits or absorbs radiation of frequency

$$f = \frac{\text{(higher energy) - (lower energy)}}{h}$$

where h is Planck's constant.

Note: Such discrete energy changes would appear as line spectra, where each line represents a specific energy change.

Postulate 3. The stationary states of an atom are those for which the angular momentum, mvr, of the atom is an integral multiple of h/2.

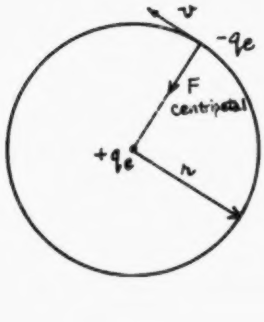
Consequences of Postulate 1:

Coulombic force = centripetal force

$$C\,\frac{q_e^2}{r^2} = \frac{mv^2}{r}$$

or, $$C\,q_e^2 = mv^2 r\,.$$

Here the unknowns are v and r.



Invoke Postulate 3.

$$mvr = n\,\frac{h}{2\pi}\,.$$

With these two equations, the unknowns can be determined. From Postulate 3,

$$v = \frac{nh}{2\pi mr}$$

and thus,

$$Cq_e^2 = m\,\frac{n^2h^2}{4\pi^2m^2r^2}\,r = \frac{n^2h^2}{4\pi^2mr}\,.$$

Solving for r_n:

$$r_n = \frac{n^2h^2}{4\pi^2mCq_e^2}\,.$$

In this equation all the factors on the right are known so r can be computed. At this point we have a check-point for the theory. The diameter of hydrogen was known from kinetic data to be on the order of 1 Å. The above equation gives a value for the radius of 0.529 Å. This is a very encouraging result

Defining r_1 as the value of r when the integer n equals 1, we can write,

$$r_1 = \frac{h^2}{4\pi^2mCq_e^2}\,.$$

The radius for an arbitrary orbit is $r_1 = n^2r_1$. This result means that only orbits with certain radii are permitted. Since $r_1 = 0.529$ Å, we have

$$r_2 = 4\ (0.529\ \text{Å}) = 2.12\ \text{Å}$$

$$r_3 = 9\ (0.529\ \text{Å}) = 4.76\ \text{Å}$$

etc. On this model no other intermediate orbits, such as r = 3.5 Å, can exist.

Now the velocity of the electron can be determined.

$$\begin{aligned} v &= \frac{nh}{2\pi mr} \\ &= \frac{nh}{2\pi m} \times \frac{4\pi^2mCq_e^2}{n^2h^2} \\ &= \frac{2\pi Cq_e^2}{nh} \end{aligned}$$

or,

$$v_1 = \frac{2\pi Cq_e^2}{h} \text{ and } v_n = \frac{1}{n}\,v_1\,.$$

Again the factors on the right hand side are known so that v can be computed.

With the velocity known, the energy can be determined. The total energy is the sum of the kinetic energy and the electrostatic potential energy, or

$$E = \frac{1}{2}mv^2 + \left(-C\frac{q_e^2}{r}\right) .$$

Substituting our derived expression for r and v, we can write:

$$E = \frac{1}{2}m\,\frac{4\pi^2C^2q_e^4}{n^2h^2} - Cq_e^2\,\frac{4\pi^2mCq_e^2}{n^2h^2}$$

$$= \frac{2\pi^2mC^2q_e^4}{n^2h^2} - \frac{4\pi^2mC^2q_e^4}{n^2h^2}$$

$$= \frac{2\pi^2mC^2q_e^4}{n^2h^2} .$$

Or we can write $E_n = \frac{1}{n^2}E_1$, where E_1 can be computed from the known quantities to be equal to -13.6 eV. This was another experimental check-point of the Bohr theory, for the energy needed to remove the electron from the hydrogen atom was known to be approximately 13.6 eV. Note that the atom can exist only in certain energy states; namely,

$E_1 = -13.6$ eV

$E_2 = -\frac{1}{4}\,13.6$ eV $= -3.40$ eV

etc. In terms of this model, there is no energy of -10.0 eV.

A gravitational analogy might be instructive at this point in the development. A gravitational "well" is pictured below. At "ground level" the potential energy is, by definition, zero. A stone can exist in stable positions (stationary states) at certain potential energies below the ground level which is defined as zero. Thus, energy must be absorbed for the stone to undergo a transition from the -10 level to the -4 level.

Sec. 19.6 The Bohr theory: the spectral series of hydrogen

Bohr's first postulate is an ad hoc postulate stating that atoms do not normally radiate energy. Yet, atoms do radiate energy in a very specific way as they give rise to spectral lines. This is where Bohr's second postulate becomes important. Bohr's second postulate states that a photon is emitted by an atom when a change from a particular high-energy state to a particular low-energy state is made. A photon is absorbed when a change is made from a low-energy state to a high-energy state. Bohr's second postulate can be written in formula form as follows:

$$hf = E_h - E_\ell .$$

The subscripts h and ℓ stand for higher and lower, where, in dealing with negative energies, the lower energies have the larger magnitudes (see Fig. 19.11).

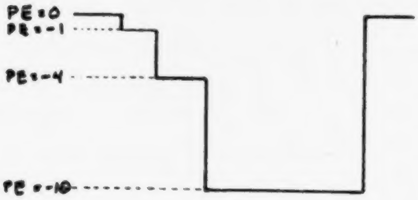


Fig. 19.11 A gravitational analogy of the Bohr atom

Let us write Bohr's second postulate in terms of the results of the last section.

$$hf = E_h - E_\ell$$

$$= \frac{E_1}{n_h^2} - \frac{E_1}{n_\ell^2}$$

$$= E_1\left(\frac{1}{n_h^2} - \frac{1}{n_\ell^2}\right)$$

or

$$hf = -\left[\frac{2\pi^2mC^2q_e^4}{h^2}\right]\left(\frac{1}{n_h^2} - \frac{1}{n_\ell^2}\right).$$

All the factors in the first parentheses are known constants. The appearance of the integers in the second parentheses is reminiscent of the Balmer formula. Balmer's formula was however written in terms of wavelength rather than frequency. When the last equation is rewritten in terms of wavelength, it can be compared to the Balmer formula. (Note the change of algebraic sign.)

$$\frac{1}{\lambda} = \frac{f}{c} = \frac{2\pi^2mC^2q_e^4}{ch^3}\left(\frac{1}{n_\ell^2} - \frac{1}{n_h^2}\right)$$

where the signs have been changed to make the first term positive.

The Balmer formula was written in terms of an empirical constant, the Rydberg constant, whose value is 109,677.58 cm^{-1}. The question is, what is the value of the fraction in the derived formula above which is made up of known constants? Does it equal 109,677.58 cm^{-1}? It does! This was indeed a triumph for the Bohr theory.

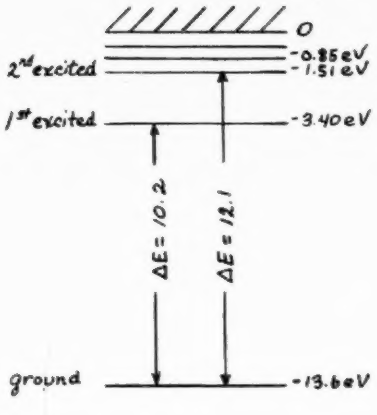
The Balmer series of spectral lines occurs when n = 2 and $n_h = 3,4...$ Thus, we can derive the Balmer formula in its entirety. In deriving it we understand it. Now the origin of spectral lines can be explained and understood. The mechanism at the atomic level responsible for the production of spectral lines is known. One of the main goals of an atomic model has been reached.

Sec. 19.7 Stationary states of atoms: the Franck-Hertz experiment

The Bohr theory predicts that the energy of an atomic electron is quantized according to the relation $E_n = E_1/n^2$. A direct proof of the existence of discrete energy states in atoms and a confirmation of Bohr's view of the origin of emission and absorption spectra is provided by the experiment of Franck and Hertz.

The experiment can be understood in terms of energy principles. The energy states of the hydrogen atom are represented in the energy-level diagram below. The first excited state is 10.2 eV above the ground state and the second excited state is 12.1 eV above the ground state. According to the Bohr theory of the hydrogen atom, its energies are precisely defined.

Franck and Hertz studied the collisions between electrons and heavy monatomic atoms. However, since we have already solved for the energies of the hydrogen atom, we shall use it as an example. (Hydrogen is diatomic and the Franck-Hertz experiment can only be done with difficulty when hydrogen is used.) The energy of the bombarding electrons can be controlled by controlling the potential through which the electron "falls."



Let us establish a potential difference of 5 volts across the tube. With a potential difference of 5 volts, an electron can obtain a kinetic energy of 5 eV. If a 5-eV electron collides with a hydrogen atom, an elastic collision occurs; that is, the electron has essentially as much energy after the collision as before the collision. The same would be true if the electron had any energies of less than 10.2 eV.

However, when the electron has an energy of 10.2 eV, a new result appears. The electron no longer collides elastically, but inelastically; that is, it loses energy in the collision. In fact, it loses all its energy! Now 10.2 eV is just the energy difference between the first excited state and the ground state.

An atom can also gain energy by photon absorption and lose energy by photon emission. In fact, in the Franck-Hertz experiment, after the atom gains energy by collisional excitation, it loses energy by photon emission. The emitted photon is the signal that excitation by electron collision has occurred.

When the electron energy reaches a value of 12.1 eV, again the electron loses all its energy in the collision. The difference between the second excited state and the ground state is 12.1 eV!

Thus the Franck-Hertz experiment is a demonstration that excitation of atoms by collision is governed by the Bohr quantization of energy.

Sec. 19.8 The periodic table of the elements

The Bohr theory provided a model of the atom which can be correlated with the periodic table and the periodic law. In the Bohr model, electrons move in well defined orbits. The chemical and physical properties of an element depend upon the arrangement of the electrons about the nucleus. To account for the periodic table and the periodic law, we must be able to determine the arrangement of electrons in the atom and show that the chemical and physical properties follow from this arrangement. This could not be done in a rigorous fashion until after the advent of quantum mechanics, yet the Bohr model was suggestive of the coming solution.

The electrons in an atom can be regarded as grouped into shells and subshells. Each shell and subshell has a fixed capacity for electrons; that is, no more than a certain number of electrons can be accommodated. The chemical and physical properties are related to the relative "emptiness" or "fullness" of the shells. For example, the inert gases have completely full shells. Thus, full shells can be associated with stability.

In Bohr's periodic table pictured in Fig. 19.12, the inert gases form the turning points in the progression of the elements. The elements just before the inert gases are short one electron of having a full shell. These elements, short one electron, are the halogens. Thus, the halogens are found to be prone to react chemically with elements from which an electron can be captured, hence filling to capacity their shells. Likewise, the elements just after the inert gases, the alkali family, have one electron in excess of a full shell. Thus, the alkali metals are prone to react chemically with elements to which an electron can be given, hence leaving them with a filled shell. One could predict that the halogens and the alkali metals are ideally suited to react with each other.

In this manner the periodic table and periodic law are explained. Thus, the Bohr model was instrumental in reaching another of the goals we set up for any atomic model.

Sec. 19.9 The failure of the Bohr theory and the state of atomic theory in the early 1920's

The Bohr model has been eminently successful. The goals established earlier have been reached. In review, the nuclear model of the atom was rendered stable by one major ad hoc postulate. The origin and mechanism of spectral lines was explained by the theory. The periodic law was given a basis in atomic structure.

In addition to these obvious successes, there was a more subtle one; namely, the Bohr theory left its indelible mark on physics. Bohr's emission and absorption of photons between stationary states remains predominant in the minds of spectroscopists. Bohr's model set the stage for further work. His quantization of angular momentum and energy was the beginning of a vital part of quantum mechanics.

Yet, for all its successes, it did not survive. Many questions which the theory was unable even to begin to answer concerned the intensities of spectral lines and the effects of a magnetic field on an atomic spectrum. To quote a contemporary physicist:

> The instant relief which Bohr's theory provided in sorting out the hopeless muddle of spectroscopic data, down to the wavelengths of the x rays, at first overshadowed all other considerations. Then, in the following years the strange emptiness of this ingenious and successful model began to impress itself on the minds of the physicists.

One of the principal reasons that the Bohr theory did not survive was that it represents a hybrid between classical and quantum ideas. It was not until new quantum ideas replaced classical ideas that these questions were answered.

Sec. 20.1 Some results of relativity theory

The theory of relativity brought about a revolution in the thinking of scientist and nonscientist alike. Because of its revolutionary nature, it has been regarded as an abstract theory, extremely difficult to understand. This is not the case. The theory of relativity is not abstract (at least not the special theory). The difficulty in understanding relativity occurs because some of our most basic concepts concerning space and time have to be re-examined and modified.

In Units 1 and 2 we have seen the difficulty people have had in accepting new modes of explanation. The idea that an object in violent motion tends to remain in motion was completely foreign to the Aristotelian natural philosopher. Another idea that was difficult to assimilate was the earth's daily rotation and annual revolution about the sun as assumed in the Copernican system. The concepts of inertia and axial rotation were only slowly accepted into the mainstream of man's thought. Likewise, some time will be required for the concept of relativity to become a natural part of man's thinking processes.

In Newtonian physics, physical phenomena are described in terms of position coordinates and momenta. When the description of a physical system is complete, it is sometimes desirable to express the state of the system in terms of a reference frame moving relative to the original reference frame. The results of transforming our description of events from one reference frame to another, and the form of physical law in arbitrary reference frames is the concern of relativity theory.

The measurement of the speed of light has been of interest to scientists since the time of Galileo. Galileo's attempts to measure the speed of light led him to believe that the propagation of light is instantaneous. As he wrote in his Two New Sciences:

> Everyday experience shows that the

> propagation of light is instantaneous; for when we see a piece of artillery fired at a great distance, the flash reaches our eyes without lapse of time; but the sound reaches the ear only after a noticeable interval.

The first successful experiment to measure the speed of light was concluded in 1675 by a Danish astronomer, Olaf Römer. From a study of the periods of Jupiter's satellites, Römer concluded that the speed of light was finite and was in the neighborhood of 200,000 miles/sec.

A great synthesis occurred when Maxwell showed that electromagnetic waves should propagate at the speed of light. This result suggested that light was electromagnetic in nature. However, Maxwell's equations gave the speed of light as a constant. Questions soon arose about the frame of reference to which Maxwell's value referred. The Michelson-Morley experiment was an attempt to answer these questions by isolating some absolute frame of reference. However, the Michelson-Morley experiment failed to do so.

Much confusion followed the failure to establish an absolute frame of reference and the way was cleared for Einstein's ideas. Einstein's first postulate, that the speed of light is a constant for all observers, "solved" the problem. With this postulate, an absolute frame of reference was no longer necessary.

The first and second postulate together form a basis from which many deductions can be made. Among these deductions are the equivalence of mass and energy and the velocity dependence of mass, length and time.

The theory of relativity forced a reexamination of some of the fundamental physical concepts. The twentieth century has seen two such periods of reassessment. The second period of reexamination took place after the development of the quantum theory. These two periods of introspection have taught physicists to be more critical of the common-sense notions that tend to dominate their thinking. (One might recall the common-sense appeal of the geocentric system.) Werner Heisenberg, one of the chief architects of quantum theory, has emphasized this point. Speaking of relativity, he says:

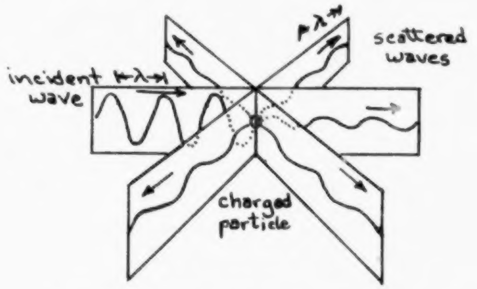
> It was the first time that scientists learned how cautious they had to be in applying the concepts of daily life to the refined experience of modern experimental science...This warning later proved extremely useful in the development of modern physics, and it would certainly have been still more difficult to understand quantum theory had not the success of the theory of relativity warned the physicists against the uncritical use of concepts taken from daily life or from classical physics.

Sec. 20.2 Particle-like behavior of radiation

In Chapter 18 we studied the photoelectric effect in which the theoretical treatment assumed a particle nature of light. In the photoelectric effect, a photon loses all of its energy in ejecting a bound electron. The initial energy of the photon appears as the binding energy (work function) plus the kinetic energy of the photoelectron.

A second example where light must be treated as corpuscular is known as the Compton effect. In this example, the photon collides with an electron and is scattered. Unlike the previous example, the photon carries some energy away from the collision. This scattering experiment clearly distinguishes between the wave and particle models of light.

When light (treated as a wave) is incident upon a charged particle, it would be scattered in all directions, as the figure below shows. The incident wave sets the charge into oscillation; the oscillating charge in turn radiates electromagnetic radiant energy in all directions. However, as the figure shows, there is no change in wavelength. (There is a change in the amplitude since the energy of the incident wave is being spread in all directions.)



Early scattering experiments showed, however, that the scattered radiation was less penetrating and seemed to have a longer wavelength than did the incident radiation. This observation contradicts the prediction based on the wave theory of light. However, when light is considered as photons, i.e., particles, one can predict that the scattered radiation should have a longer wavelength than the initial radiation, as is observed.

If we now endow our light particle with all the properties of a particle, namely, energy and momentum, we can derive the change in the wavelength of a scattered photon. The result of this derivation is as follows:

$$\Delta\lambda = \lambda - \lambda_o = \frac{h}{m_o c}(1 - \cos\phi)$$

where $\Delta\lambda$ is the change in wavelength, λ_o is the wavelength of the incident photon, λ is the wavelength of the scattered photon, and ϕ is the angle through which the photon is scattered.

The derivation of this formula is very tedious algebraically; however, the argument in terms of physical concepts is very simple and is based upon the conservation laws studied in Unit 3. Let the frequency of the incident photons be f_o and the frequency of the scattered photons be f. Then, if energy is conserved, we must have

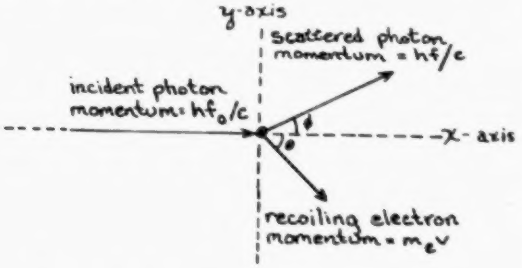
$$hf_o = hf + 1/2\, m_o v^2,$$

where m_o is the rest mass of the electron. If momentum, $m_o v$ (a vector quantity) is also conserved, we have two equations representing two components. Thus, if momentum is conserved we may write:

$$\text{x-component: } \frac{hf_o}{c} = \frac{hf}{c}\cos\phi + m_o v\cos\theta$$

$$\text{y-component: } 0 = \frac{hf}{c}\sin\phi - m_o v\sin\theta$$

where the angles are illustrated in the figure below.


y-axis
Scattered photon momentum = hf/c
incident photon momentum = hf$_o$/c
ϕ
θ
x-axis
recoiling electron momentum = m$_e$v

For a particular scattering angle ϕ, we have three equations and three unknowns: f, v and θ. We can solve these equations simultaneously for any of the unknowns. For comparison with experiment however, we are interested in the shift in the wavelength. Remembering that $\lambda_o = c/f_o$ and $\lambda = c/f$, we can rewrite the three equations in terms of $\Delta\lambda$. The result is the same as that given above and in the text.

With this equation, we can solve for $\Delta\lambda$ for any scattering angle ϕ. When we insert the known values for h, m_o and c, we get:

$$\Delta\lambda = (0.0242)\,(1 - \cos\phi)\ \text{Å}.$$

When $\phi = 90°$, $\cos\phi = 0$ and $\Delta\lambda = 0.0242$ Å. When visible light, with a wavelength on the order of 5000 Å, is used, the per cent change in the wavelength of the scattered light is too small to detect. However, when x rays, with a wavelength of approximately 1 Å, are used, the per cent change in the wavelength of the scattered x rays is observable. This change in the wavelength is known as the Compton effect.

Sec. 20.3 Wave-like behavior of matter

Apparently everything in the physical world can be classified into one of two categories, matter and radiation, a symmetry which impressed Louis de Broglie. He was also impressed by the dual nature of radiation. These considerations led de Broglie to question the particle nature of matter. Why, he asked, should not the same dualism which characterizes the basic quanta of radiation, also characterize the basic quanta of matter? That is, should particles also have wave characteristics? He developed his ideas and presented them in the form of a doctoral dissertation.

As we have stated them, de Broglie's ideas sound highly speculative. However, de Broglie did much more. First, he gave a definitive relation equating the wave properties with the particle properties (page 108). Second, he was struck by the fact that the stability condition for atomic orbits, namely the quantization of angular momentum, introduced integers. With his postulate he was able to deduce the quantum condition of Bohr which we studied in Chapter 19. Thus, the wave nature of matter gave a basis for understanding the Bohr atom; that is, those orbits are allowed which represent an integral number of wavelengths.

Of course the impressive result came when the de Broglie hypothesis was verified experimentally. To demonstrate wave properties, one does a diffraction experiment which crucially depends on the wave nature. Davisson and Germer showed that electrons could be diffracted and thus demonstrated the wave nature of electrons.

Sec. 20.4 Quantum mechanics

The first step in understanding the internal structure of the atom was provided by the work of Rutherford and Bohr, discussed in Chapter 19. It might be argued that this was the most decisive step. Yet, while Bohr's theory was stimulating, it lacked the breadth to accommodate all the facts and could not be structured to answer the vital questions. It became obvious that a new, more fundamental approach was needed. Bohr himself was one of the leaders in the quest for a deeper understanding.

We have seen that the first big breakthrough came with the de Broglie hypothesis. On the basis of this hypothesis, Schrödinger developed a mathematical formalism which came to be known as wave mechanics. The mathematical representation of waves has deep roots in classical physics. Water waves, sound waves and radio waves are significantly different. The first is a transverse mechanical wave, the second a longitudinal mechanical wave, and the third an electromagnetic wave. However, in spite of their differences, the same type of mathematical equation will describe all three. Schrödinger, viewing the electron as a wave, wrote an equation which is very similar in mathematical form to the equations that describe water waves, sound waves and light waves. His equation is written below. (This equation is not necessarily to be shown to the student; however, there may be some students whose curiosity has been aroused and would like to see this famous equation):

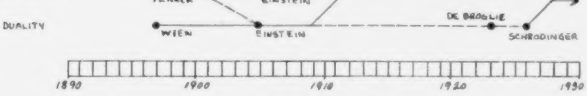
$$-\frac{h^2}{8\pi^2 m}\left(\frac{\partial^2\psi}{\partial x^2}+\frac{\partial^2\psi}{\partial y^2}+\frac{\partial^2\psi}{\partial z^2}\right)$$

$$-\left(\frac{e^2}{\sqrt{x^2+y^2+z^2}}\right)(\psi) = E\psi$$

With the Schrödinger equation, properties of the hydrogen atom can be computed which include not only the results of Bohr, but many more besides.

Heisenberg developed an alternate approach to quantum mechanics. His approach grew more directly from the experimental data of atomic spectra. The two approaches, Schrödinger's and Heisenberg's, were ultimately shown by Dirac to be equivalent.

The various lines of investigation which converged to form the impetus for the quantum mechanical revolution are illustrated below.



Sec. 20.5 Quantum mechanics—the uncertainty principle

One of the consequences of the wave-particle dualism is the uncertainty principle, first enunciated in 1927 by Heisenberg. This principle sets a fundamental limit on the ultimate precision with which we can simultaneously know both the position of a particle and its momentum.

In Newtonian mechanics, even as modified by special relativity, the position and momentum relative to a given frame of reference can be both exactly and simultaneously defined. Then, with Newton's laws of motion, all future motions of the particle can be accurately determined. This formed the basis for a philosophy of mechanistic determinism. One of the most ardent disciples of this philosophy was Laplace, a French mathematician. He articulated the essence of this philosophy as follows:

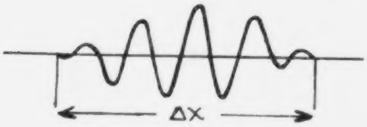
> An intellect which at any given moment knew all the forces that animate nature and the mutual positions of the beings that compose it, if this intellect were vast enough to submit its data to analysis, could condense into a single formula the movement of the greatest bodies of the universe and that of the lightest atom: for such an intellect nothing could be uncertain; and the future just like the past would be present before its eyes.

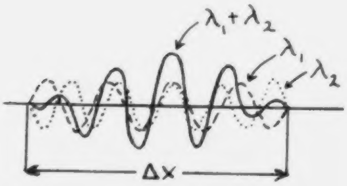
This statement represents quite a contrast with current physical thought.

The conclusion of current physical thought is that we are unable to know the present exactly, hence we cannot know the future exactly. For example, if we attempt to measure precisely the position of an electron, we lose all knowledge of its momentum. The very act of measurement changes the position or motion of the particle (as several examples on page 119 of the text suggest), with the result that its future positions cannot be precisely predicted.

It is quite simple to see how the uncertainty principle arises from the de Broglie hypothesis, as developed on page 118. This hypothesis relates the wavelength of a particle with its momentum. Therefore, if we know the momentum exactly, we also know the wavelength exactly. But any wave with an exact wavelength (such as a sine wave) is a continuous wave having infinite extent. If, then, a particle is represented by a wave having infinite extent, its position is completely undefined. We conclude, therefore, that precise momentum (exact wavelength) implies an undefined position.

A localized wave, in contrast to an infinite wave, can be built up from a superposition of sine waves. The figure below shows such a localized wave. A particle, represented by such a wave packet, can be located within an uncertainty of Δx. The synthesis of such a wave packet from two sine waves of wavelengths λ_1 and λ_2 is shown below. The question now is, what is the wavelength (momentum) of the particle? The superposition of two waves of different wavelengths is required to give the particle some localizability. Thus, in localizing the particle we have lost our precise knowledge of its momentum (wavelength). And so it goes. We purchase knowledge of the one at the price of uncertainty of the other. Complete knowledge at the price of complete ignorance.





Further discussion of the uncertainty principle appears in the January 1958 Scientific American in an article by George Gamow, "The Principle of Uncertainty."

Sec. 20.6 Quantum mechanics—probability interpretation

In Born's interpretation, the square of the amplitude of the de Broglie wave is related to the probability of finding a particle. A particle is most likely to be observed in regions of large amplitude.

How does one find these amplitudes? They are found by solving the Schrödinger equation. As was pointed out in Section 20.4, Schrödinger's equation describes the motions of electrons. We now see that Schrödinger's equation gives us an amplitude, the square of which is a probability. Thus, we can compute where an electron is most likely to be.

These considerations have forced physicists and chemists to change their conception of the atom. The Bohr model of the hydrogen atom pictured the electron traveling in prescribed "planetary" orbits about the nucleus. Such a view is inconsistent with the intrinsic nature of the electron, that is, its dual nature. As the figures on pages 119 and 120 of the text indicate, the electron having an energy of -13.6 eV can be found anywhere. One is most likely, however, to find the -13.6 eV electron at a distance of 0.529 Å from the nucleus, which corresponds to the first Bohr orbit. Thus, that which is most probable on the submicroscopic level corresponds to that which is observed on the macroscopic level.

This relation between the microscopic and the macroscopic is reminiscent of the kinetic theory of gases. In kinetic theory, the average value of dynamic variables on the molecular level is related to observable parameters. Our ability to compute the values of these observables applies only to those systems made up of large numbers of particles.

In a similar fashion, we are able to compute the values of observables in quantum theory because we are dealing with extremely large numbers of particles. With large numbers of particles, that which is highly probable for one, overwhelmingly determines the macroscopic behavior of the system of particles.

Summary Sheet—Transparencies

T35 Periodic Table

The modern long form of the periodic table is presented with various overlays highlighting chemical families and other pertinent groupings.

T36 Photoelectric Mechanism

Schematic drawings of a photoelectric tube and circuit show the procedure for measuring stopping voltage.

T37 Photoelectric Equation

A sliding mask on a plot of maximum kinetic energy of photoelectrons versus frequency of photons permits the derivation of Einstein's photoelectric equation.

T38 Alpha Scattering

A two-part transparency showing a diagrammatic sketch of the Rutherford scattering experiment, and potential hill diagrams for the Thomson and Rutherford atomic models.

T39 Energy Levels—Bohr Theory

A two-part transparency showing the Bohr orbits and energy levels for hydrogen. Illustrates production of Lyman and Paschen Series in general and Balmer Series in detail.

Film Loops

L46 Production of Sodium by Electrolysis

Davy's classical experiment in which molten NaOH is electrolyzed to form metallic sodium.

L47 Thomson Model of the Atom

Small magnets floating on the surface of water are aligned into various patterns by a radial magnetic field. The apparatus, a model of a model, was described by Thomson.

L48 Rutherford Scattering

A computer-animated film, in which projectiles are fired toward a nucleus which exerts an inverse-square repulsive force.

16MM Films

F35 Definite and Multiple Proportions

Here is the evidence on which Dalton based his conviction that matter came in natural units, atoms; the chemical laws of definite proportions demonstrated by electrolysis and recombination of water; and multiple proportions by the quantitive decomposition of N_2O, NO and NO_2. 30 minutes, Modern Learning Aids.

F36 Elements, Compounds and Mixtures (color)

A discussion of the difference between elements, compounds and mixtures, showing how a mixture can be separated by physical means. Demonstrates how a compound can be made and then taken apart by chemical methods, with identification of components by means of their physical properties such as melting point, boiling point, solubility, color, etc. 33 minutes, Modern Learning Aids.

F37 Counting Electrical Charges in Motion

This film shows how an electrolysis experiment enables us to count the number of elementary charges passing through an electric circuit in a given time and thus calibrate an ammeter. Demonstrates the random nature of motion of elementary charges, with a current of only a few charges per second. 22 minutes, Modern Learning Aids.

F38 Millikan Experiment

Simplified Millikan experiment described in the text is photographed through the microscope. Standard spheres are substituted for oil drops; an analysis of the charge related to the velocity of the sphere across the field of view of microscope emphasizes the evidence that charge comes in natural units that are all alike; numerous changes of charge are shown, produced by x rays, with the measurements clearly seen by the audience. 30 minutes, Modern Learning Aids.

F39 Photoelectric Effect (color)

Qualitative demonstrations of the photoelectric effect are shown using the sun and a carbon arc as sources. A quantitative experiment is performed measuring the kinetic energy of the photoelectrons emitted from a potassium surface. The data is interpreted in a careful analysis. 28 minutes, Modern Learning Aids.

F40 The Structure of Atoms

This film provides the experimental evidence for our basic concepts concerning the structure of the atom. An experiment similar to Rutherford's historic alpha-particle scattering demonstrations shows that atoms have dense, positively charged nuclei. Another fundamental experiment shows the charge on the electron and the ratio of charge to mass. 12½ minutes—B&W—code 612010; 12½ minutes—color—code 612022. McGraw-Hill.

F41 Rutherford Atom

A cloud chamber and gold foil in a simple alpha-particle scattering experiment to illustrate the historic Rutherford experiment which led to the nuclear model of the atom. Behavior of alpha particles clarified by use of large-scale models illustrating the nuclear atom and Coulomb scattering. 40 minutes, Modern Learning Aids.

F42 A New Reality

Traces the discovery of the structure of the atom and emphasizes the work of the Danish physicist, Niels Bohr. The story begins at the Institute for Theoretical Physics in Denmark, where experts from all parts of the world study and experiment with the atom. Man has devised means of visualizing the sub-microscopic structure of molecules and by advanced electronic equipment has gained an understanding of the character of the atom. We see how one element can be converted to another by atomic bombardment which changes the number of protons in the nucleus. Other demonstrations using light waves establish color measurement in terms of energy. Also illustrated are proofs that the electron components of the atom are both particles and wave energies. The modern concept of the atom is basically that determined by Niels Bohr, and its implications reach into the realms of biology, psychology and philosophy. Produced by Statens Filmcentral and Laterna Films, Denmark, and OECD and Sponsored by the Carlsberg Foundation in Association with the International Council for Educational Films. 1965 release (3/2 IFB 394 color) 51 minutes, International Film Bureau Inc.

F43 Franck-Hertz Experiment

A stream of electrons is accelerated through mercury vapor, and it is shown that the kinetic energy of the electrons is transferred to the mercury atoms only in discrete packets of energy. The association of the quantum of energy with a line in the spectrum of mercury is established. The experiment retraced in the film was one of the earliest indications of the existence of internal energy states within the atom. 30 minutes, Modern Learning Aids.

F44 Interference of Photons

An experiment in which light exhibits both particle and wave characteristics. A very dim light source, a double slit and a photomultiplier are used in such a way that less than one photon (on the average) is in the apparatus at any given time. Characteristic interference pattern is painted out by many individual photons hitting at places consistent with the interference pattern. Implications of this are discussed. 13 minutes, Modern Learning Aids.

F45 Matter Waves

The film presents a modern version of the original experiment which showed the wave behavior of the electron. The student sees electron diffraction patterns on a fluorescent screen. The patterns are understandable in terms of wave behavior. Alan Holden presents an optical analogue showing almost identical patterns. The electron diffraction experiments of G. P. Thomson are described by Holden, who also presents brief evidence for the wave behavior of other particles such as neutrons and helium atoms. 28 minutes, Modern Learning Aids.

F46 Light: Wave and Quantum Theories

This film clearly and simply demonstrates the accepted theory of light as consisting of both a wave motion and of discrete bundles, or quanta of energy. Young's double-slit experiment is performed to show the wave character of light, while the photoelectric effect indicates that light consists of energy quanta. The Compton effect and other major experiments associated with the present theory are shown in laboratory demonstrations and in animation. B & W and Color. 13½ minutes, Coronet Films.

Film Sources

Modern Learning Aids
1212 Avenue of the Americas
New York, New York 10036

McGraw-Hill Book Company
Text Film Department
330 West 42nd Street
New York, New York

Coronet
42 Midland Road
Roslyn Heights
New York

International Film Bureau Inc.
332 South Michigan Avenue
Chicago, Illinois 60604

A more complete film source reference is given in the Unit 2 Teacher Guide, pages 73-77.

Reader Articles

1. FAILURE AND SUCCESS
Charles Percy Snow
1958

This author describes the frustrations and joy that can accompany a scientific discovery. The book is based on Snow's early experiences as a physical chemist.

2. STRUCTURE, SUBSTRUCTURE, SUPERSTRUCTURE
Cyril Stanley Smith
1965

The seemingly arbitrary forms of soap bubbles, crystals, bee hives, plant cells, and other aggregates of structures are subject to laws of physics and hence follow predictable patterns.

3. THE ISLAND OF RESEARCH
Ernest Harburg
1966

4. THE "THOMSON" ATOM
J. J. Thomson
1966

J. J. Thomson's atomic model was an early attempt to explain the emission spectrum of elements and the presence of the newly-discovered electrons in atoms. In this century more successful atomic theories were devised by Rutherford, Bohr and Schrödinger.

5. EINSTEIN
Leopold Infeld
1941

A noted Polish theoretical physicist and co-worker of Albert Einstein takes us into the study of the great twentieth-century physicist.

6. MR. TOMPKINS AND SIMULTANEITY
George Gamow
1965

Mr. Tompkins takes a holiday trip in a physically possible science-fiction land. In solving a murder case there he learns the meaning of the concept of simultaneity in the theory of relativity.

7. MATHEMATICS AND RELATIVITY
Eric M. Rogers
1960

Rogers, a noted physics teacher, introduces the fundamental concepts of the theory of relativity and illustrates the relation of mathematics to physics.

8. RELATIVITY
Anonymous
1955

9. PARABLE OF THE SURVEYORS
Edwin F. Taylor and John Archibald Wheeler
1966

Invariance is central to the theory of relativity as to all modern physics. The story told here introduces many of the important fundamental concepts of relativity theory.

10. OUTSIDE AND INSIDE THE ELEVATOR
Albert Einstein and Leopold Infeld
1961

The father of the general theory of relativity and his associate illustrate one of the central ideas of the theory through the commonplace experience of riding in an elevator. (Note: The initials C. S. mean "coordinate system" in this selection.)

11. EINSTEIN AND SOME CIVILIZED DISCONTENTS
Martin Klein
1965

What lessons can be learned from the life and philosophy of a "high-school drop-out" named Albert Einstein? Martin Klein, a physicist and historian of science, discusses the possibility of inadequacies in our present education policies.

12. THE TEACHER AND THE BOHR THEORY OF THE ATOM
Charles Percy Snow
1958

We visit, in this brief passage, an elementary science class hearing for the first time about the Bohr theory of the atom.

13. THE NEW LANDSCAPE OF SCIENCE
Banesh Hoffmann
1959

Educated as we are in classical physics, we may be unprepared to comprehend the world of quantum mechanics. This book tries to introduce us to this new view of the world.

14. THE EVOLUTION OF THE PHYSICIST'S PICTURE OF NATURE
Paul A. M. Dirac
1963

An account of how physical theory has developed in the past and how it might be expected to develop in the future.

15. DIRAC AND BORN
Leopold Infeld

Infeld reminisces what it was like to work at Cambridge University in England with two great, but very different, theoretical physicists.

16. I AM THIS WHOLE WORLD: ERWIN SCHRÖDINGER
    Jeremy Bernstein
    1961

    Erwin Schrödinger was the inventor of the basic equations of modern atomic theory. This article considers a book in which Schrödinger discusses the repercussions of the quantum theory.

17. THE FUNDAMENTAL IDEA OF WAVE MECHANICS
    Erwin Schrödinger
    1933

    A master of the physics of the atom explains how he arrived at the modern theory of the atom. This lecture is not easy, but it is worth working through.

18. THE SENTINEL
    Arthur C. Clarke
    1953

    The moon explorers make an unexpected discovery, and react in an all-too-human way in this short story by a well-known writer of science fiction.

19. THE SEA-CAPTAIN'S BOX
    John L. Synge
    1951

    A distinguished mathematical physicist, the nephew of the great Irish playwright John Millington Synge, uses an amusing allegory to discuss the nature of scientific knowledge.

20. SPACE TRAVEL: PROBLEMS OF PHYSICS AND ENGINEERING
    Harvard Project Physics Staff
    1960

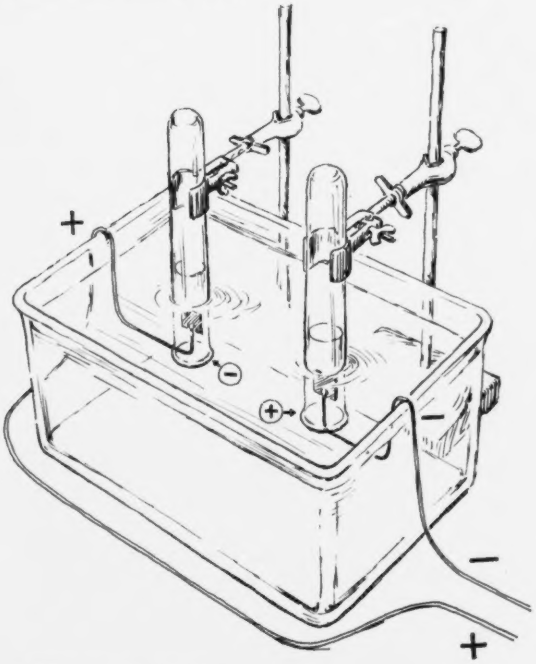
    This article, based on lectures of Edward M. Purcell, distinguishes between sound proposals and unworkable fantasies about space travel.

21. LOOKING FOR A NEW LAW
    Richard P. Feynman
    1965

    A successful theoretical physicist discusses informally in this talk the process of discovering physical theories.

D53: Electrolysis of Water

Set up the apparatus as shown in the figure. The electrodes are clamped alongside the inverted test tubes and connected either to a 6-volt battery or power supply. Any gas that forms on either of the platinum-plated spiral electrode tips will be collected when it rises and displaces water from the corresponding test tube. Since this reaction is very slow with pure water, it is necessary to add about 5 to 15 cm^3 of 6 Normal dilute sulfuric acid to the water. 6 Normal can be prepared by adding 1 volume of concentrated sulfuric acid to 5 volumes of water and stirring with a glass stirring rod.



Always pour the acid into the water. After the acid is added and stirred, a full test tube of hydrogen can be collected in about 20 minutes. Test both gases to demonstrate that hydrogen and oxygen have been produced, then fill each test tube with water from a graduated cylinder and thus measure the volume of gas produced. Calculate the ratio.

The two gases can also be collected in one test tube and ignited. There will be a violent reaction, from which it can be concluded that the gases readily combine. It is not possible to conclude that water was formed as a result of this reaction since the test tube is wet before ignition. The platinum electrodes with connecting wires are available from Macalaster Scientific and Damon Corporation (about $4.00).

D54: Charge-to-Mass Ratio for Cathode Rays

A simple demonstration of the deflection of the beam of a cathode-ray tube in a magnetic field can bring together parts of Units 4 and 5, and can give quite a good value for the ratio q_e/m. In Unit 4 (Experiment 37—The Electron-Beam Tube) students saw, qualitatively, the deflection of the cathode rays in the magnetic field of a pair of permanent magnets. In Experiment 36—Currents, Magnets and Forces, they learned how to measure magnetic fields, and one group of students used the current balance to measure the vertical component of the earth's magnetic field. This field will be used to deflect an electron beam, and its measured value used to calculate q_e/m. This ratio in itself is probably not of great interest to students, but together with q_e found in the Millikan-type experiment (Experiment 40—The Measurement of Elementary Charge) it enables us to estimate m, the mass of the electron.

Equipment

cathode-ray tube (such as the 2 inch 902A—about $15.00)

power supply for CRT (for 902A—6.3 V, 1 A for heater, about 200 V dc for anode)

(permanent magnet)

(sticky centimeter tape)

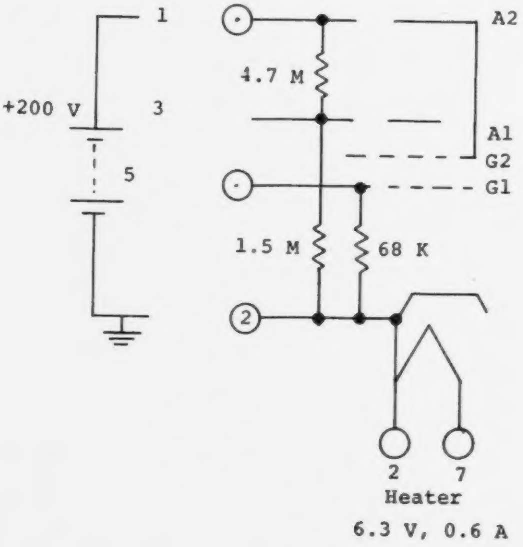


Fig. 1 Schematic for CRT 902A

Procedure

Connect the CRT to power supply. (The simplest wiring diagram for the 902A shown in Fig. 1. Use of fixed resistors will give an adequately, but not perfectly, focussed beam. To improve focus and control brightness, replace resistor R_2 by a potentiometer (about 1 Meg) and a resistor (about 0.5 Meg) in series and adjust the potentiometer for best focus. A more complete schematic, suggested by the manufacturers, is given at the end of this note (Fig. 5).

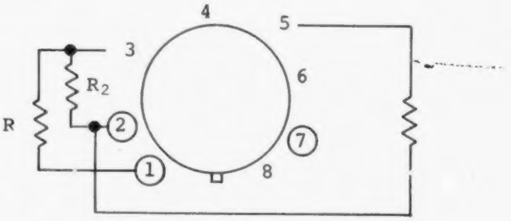

4
5
3
R2
6
R
2
7
1
8

Fig. 2 Bottom view of tube socket (medium shell octal 8-pin) with resistors. Ground pin no. 2, connect 6.3 volts between pins 2 and 7, +200 volts to pin 1.

It is important to keep the anode voltage as low as possible. With the 902A a spot is obtained with only 200 volts, but other tubes may require more. In a typical cathode-ray oscilloscope the accelerating (anode) potential is more than 1000 volts: this produces a beam of faster electrons that are correspondingly less deflected in the external magnetic field, whereas a beam of 200 volt electrons is appreciably deflected in the earth's field.

First demonstrate that the beam is affected by an applied magnetic field: bring a small permanent magnet near the tube and show the deflection of the spot.

Electric deflection can be shown by applying a potential difference of a few volts (e.g., from a six-volt battery) between the pairs of deflection plates. Note that one of the x-plates and one of the y-plates are connected internally to anode 2. Terminals 4 and 6 are connected to the other x- and y-plates. With the wiring suggested here (Fig. 1) these terminals will be "hot"—about 200 volts above ground.

The significance of this first qualitative part of the demonstration is discussed on pp. 40-41 of Unit 5. The fact that the cathode ray is deflected by magnetic (and electric) fields suggests that it consists of charged particles. That the beam remains narrow and well-defined suggests that all the particles are identical—if they were not we would expect some to be deflected more than others, which would cause the beam to spread out. J. J. Thomson went further than this and showed that the deflection was independent of the cathode material and of the residual gas in the tube.

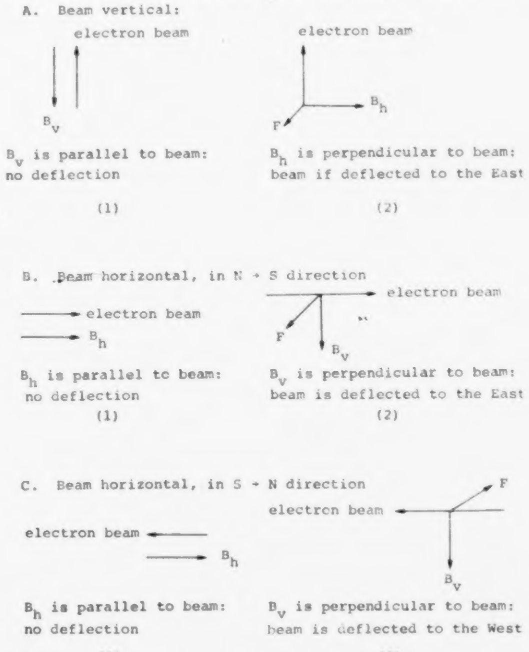

A. Beam vertical:
electron beam
B_v
B_v is parallel to beam: no deflection
(1)
electron beam
B_h
F
B_h is perpendicular to beam: beam if deflected to the East
(2)
B. Beam horizontal, in N → S direction
electron beam
B_h
B_h is parallel to beam: no deflection
(1)
electron beam
F
B_v
B_v is perpendicular to beam: beam is deflected to the East
(2)
C. Beam horizontal, in S → N direction
electron beam
B_h
B_h is parallel to beam: no deflection
(1)
F
electron beam
B_v
B_v is perpendicular to beam: beam is deflected to the West
(2)

Fig. 3

To get the effect of the vertical component of the field measure the total displacement of the spot (how much it shifts when the tube is moved from the N-S to the S-N orientation) and divide by two. A question for students: in what direction must the beam be oriented to be undeflected by the earth's field?

To make any quantitative interpretations of the deflection we must be able to measure the magnetic field. The field due to a permanent magnet is unsatisfactory for this purpose because it is non-uniform. Fortunately the (vertical component of the) earth's magnetic field is strong enough to deflect the beam appreciably; it is certainly uniform over the length of the tube; and it can be measured with the Project Physics current balance (Experiment 36, group C).

Set up the tube horizontally along a N-S line. Mark the position of the spot on the screen. Turn the tube through 180°, so that the electron beam is still horizontal, but is in the opposite direction (S-N). Mark the spot again and measure the horizontal displacement from its first position.* (A strip of sticky centimeter tape stuck on the tube face is useful.) Ignore any vertical displacement, which at most should be small.

Use a voltmeter to record the anode potential.

Theory

Electrons are emitted from the hot cathode, are accelerated by a positive potential on the anode. After passing through a hole in the anode they move with constant velocity till they strike the luminescent screen.

The kinetic energy gained by an electron due to acceleration through a potential difference of V volts is

$$\tfrac{1}{2}mv^2 = vq_e \qquad (1)$$

where q_e is the charge of an electron.

After the electrons have passed through the anode hole they can be deflected by electric and magnetic fields. The magnitude of the force (F) on an electron moving with velocity (v) in a magnetic field (B) is

$$F = Bq_e v.$$

The force is perpendicular to the directions of both v and B. In this case v is horizontal, and B vertical: F will be horizontal.

A particle is moving with constant velocity v and is acted on by a force perpendicular to that velocity: the particle will move in a circle (see Unit 1).

The centripetal force needed to keep a particle of mass m moving in a circle of radius R is mv^2/R. But we already know that

$$F = Bq_e v$$

$$\frac{mv^2}{R} = Bq_e v$$

$$\frac{q_e}{m} = \frac{v}{BR}.$$

Substituting for v from (1) above,

$$\frac{q_e}{m} = \frac{(2Vq_e)^{\frac{1}{2}}}{m^{\frac{1}{2}}\,BR}$$

$$\frac{q_e}{m} = \frac{2V}{B^2R^2}.$$

To find the radius of curvature of the electron beam, R, apply Pythagoras' theorem to triangle OPQ (Fig. 3):

$$(R - x)^2 + \ell^2 = R^2$$

$$R^2 - 2Rx + x^2 + \ell^2 = R^2$$

$$2Rx = x^2 + \ell^2$$

if $x < \ell,\ x^2 << \ell^2$

and $2Rx \approx \ell^2$

$$R \approx \frac{\ell^2}{2x}.$$

(Compare the geometry of the pendulum—Student Handbook, Unit 3, p. 10-8.)

The distance ℓ for the 902A tube is 10 cm.

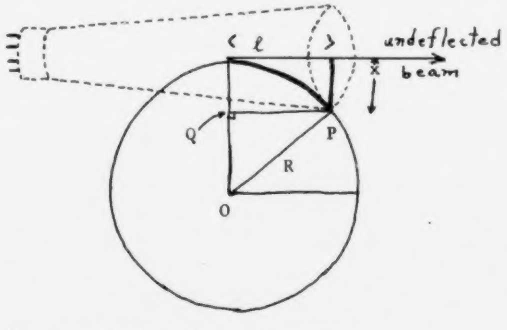

ℓ
undeflected beam
x
Q
P
R
O

Fig. 4 ℓ is length of tube from anode to screen; x is horizontal displacement of spot on the screen; R is the radius of curvature of the circle into which the electron beam is bent.

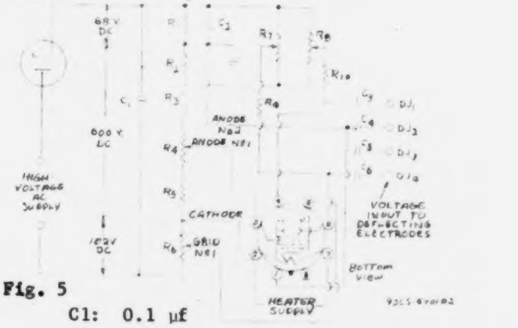

68 V DC
600 V DC
182V DC
HIGH VOLTAGE AC SUPPLY
C1
C2
R1
R2
R3
R4
R5
R6
R7
R8
R9
R10
ANODE No2
ANODE No1
CATHODE
GRID No1
C3
C4
C5
C6
DJ1
DJ2
DJ3
DJ4
VOLTAGE INPUT TO DEFLECTING ELECTRODES
BOTTOM VIEW
HEATER SUPPLY

Fig. 5

C1: 0.1 μf
C2: 1.0 μf
C3 C4 C5 C6; 0.05-μf **Blocking capacitors***
R1 R2: 1 Megohm
R3: 1.3 Megohms

*When cathode is grounded, capacitors should have high voltage rating; when anode No. 2 is grounded, they may have low voltage rating. For dc amplifier service, deflecting electrodes should be connected direct to amplifier output. In this service, it is preferable usually to remove deflecting-electrode circuit resistors to minimize loading effect on the amplifier. In order to minimize spot defocusing, it is essential that anode No. 2 be returned to a point in the amplifier system which will give the lowest possible potential difference between anode No. 2 and deflecting electrodes.

---

*Notice that the deflection is not symmetrical about the spot's position when the tube is vertical. This is because with the tube vertical the beam is deflected (to the east) by the horizontal component of the earth's field B_h [see Fig. 3A(2)].

R4: 1-Megohm Potentiometer
R5: 0.3 Megohm
R6: 0.5-Megohm Potentiometer
R7 R8: 2-Megohm Potemtiometers
R9 R10: 2 Megohms

D55: Photoelectric Effect

A simple electroscope demonstration introduces the photoelectric effect vividly in a qualitative way and displays features not shown at all in the student experiment.

The necessary equipment (Fig. 1) consists of:

1. a simple electroscope whose electrode can be surmounted or replaced by 10 cm × 10 cm pieces of zinc and copper and possibly lead and iron
2. a source of ultraviolet light such as a small mercury vapor sterilizing lamp to illuminate the metal plates
3. a sheet of ordinary glass to hold between the lamp and the metal plates
4. plastic strips or other materials for giving the electroscope positive and negative charges
5. a piece of sandpaper or steel wool for cleaning the metal plates.

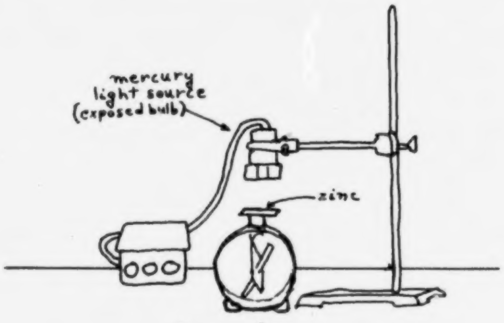


Figure 1

Procedure

The metal plates should each be scrubbed hard with sandpaper or steel wool to remove any traces of oxide. Even a few hours after cleaning it must be done again, and the plate then wiped with a clean cloth to remove any traces of sand or steel wool whose sharp points allow rapid leakage of charge.

Mount the zinc plate on the electroscope as shown in Fig. 1, and give it a negative charge. (Charging by induction often gives it a much larger charge than charging by contact. If the day is humid, it may be necessary to dry the charging materials, e.g., over a bright lamp bulb or a radiator.)

Watch the charged electroscope for a few moments to make sure its charge is not leaking at a visible rate. Then illuminate the plate with light from the mercury lamp.

The electroscope should discharge rather rapidly. A large class can see this best if the electroscope shadow is projected on the chalkboard by a bright light several feet away.

To show that the photoelectric effect is related to chemical activity and therefore to the ease with which a metal loses electrons, replace the zinc by metals of lower chemical activity such as copper, iron, and lead, cut to the same size.

To show that the effect is not caused by visible light:

(a) replace the ultraviolet lamp by an ordinary incandescent lamp. However bright, it will not drive electrons from these metals, though, of course, it will do so from more active elements such as cesium and lithium, used to coat the emitters of photoelectric cells; and

(b) shield the metal plate from the ultraviolet radiation with the glass plate. Ultraviolet does not penerate glass. The electroscope's discharge immediately stops.

None of these qualitative effects is shown in the student laboratory experiment, whose purpose is to show how the energy of the emitted electrons depends on the frequency of the light and so to demonstrate the inadequacy of the wave theory of light for photoelectric phenomena.

It is worth noting that the photoelectric effect was discovered originally by Heinrich Hertz in 1887 in the course of research that gave massive support to Maxwell's wave model of light. He found that ultraviolet light shining on the terminals of a spark gap facilitated the formation of sparks.

D56: Blackbody Radiation

As the brightness of an ordinary 150- or 200-watt incandescent lamp is varied, its color changes markedly. Students can see this easily through their pocket spectroscopes as the brightness of an incandescent lamp is varied by means of a variable transformer. Do this in a partially darkened room. Warn them to ignore the immense change in brightness and to concentrate on the changing distribution of color.

First only red is visible, faintly. Then orange, yellow, etc. are added until the bright light is almost (but not quite) white.

This demonstration leads easily to a description (not a derivation) of the radiation curves (Fig. 1, in which the intensity is plotted against wavelength for each temperature) and Planck's derivation of their formula in 1900. Note the use of these curves for finding the surface temperatures of the source of any continuous spectrum—molten steel or a distant star.

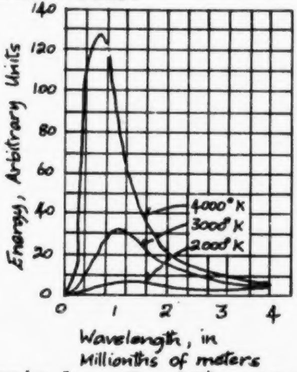


Fig. 1

Stefan's law, $E = kT^4$, and Wien's law, $\lambda_{max} = c/T$ could, perhaps, be brought out at this point too. (E is the total energy radiated by a black-body; T = absolute temperature; λ_{max} is wavelength of most intense emission, and c and k are constants.) There are not many fourth-power laws in physics.

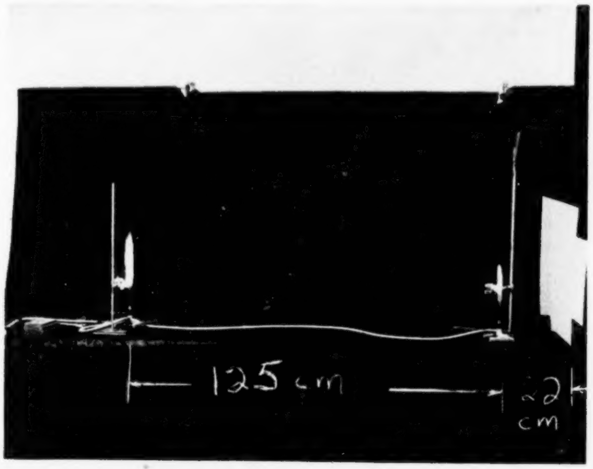
The important point to raise here is that Planck's successful explanation of the continuous spectrum energy distribution curve required for the first time the assumption that radiant energy was emitted in chunks, or "quanta" of energy hf. This was the origin of the quantum theory.

It is worth noting, too, that the next significant use of the quantum idea was by Albert Einstein twelve years later to explain the emission of photoelectrons. His photoelectric equation is described in Experiment 42*, which the students have probably already performed.

D57: Absorption

This demonstration shows the absorption of light by excited molecules. It should be done in a dark room.

Set up two bunsen burners and a white screen as shown in the figure. The distances are approximate only.



The flame that is further from the screen should be turned up higher. When the two flames are burning steadily, with luminous flames (air supply shut off) the further one will cast a shadow of the nearer burner on the screen.

Now open the air holes of both burners and introduce some sodium into each flame. Look carefully at the screen and you will see that the flame of nearer burner also casts a shadow. Some of the light emitted by excited Na atoms in the further flame is absorbed by Na atoms in the nearer flame. Adjust the brightness and distances to get maximum effect. Try not to set up air currents that will disturb the flames and make the shadows flicker excessively.

To bring home the point that it is the same monochromatic light emitted in one flame and absorbed in the other that causes the shadow, remove the sodium from the further flame and cut off the air to make it luminous again. The nearer flame casts no shadow now because nearly all the light emitted by the farther flame passes through it without being absorbed. But the flame is still bright enough to cast a shadow of the burner tube. If you remove the sodium from the nearer flame the effect will be the same: no absorption and therefore no shadow.

D58: Ionization Potential

The existence of a discrete ionization potential for each element confirms the existence of stationary states, postulated by Bohr and mathematically derived by Schrödinger. The ionization potential of argon can be demonstrated rather simply using a thyratron 884 tube.

Connect the tube as shown in Fig. 1. Notice that since the grid and the plate are connected together the tube is essentially a diode.

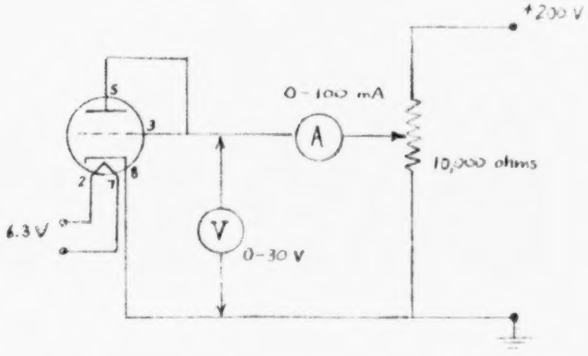


Fig. 1

A power supply with a low internal resistance, capable of delivering a high current, should be used. The variable resistance is an ordinary 5000-ohm potentiometer.

If the 5000-ohm potentiometer is not available, the voltage can be varied instead by running the power supply from a variable transformer. Of course the filament of the 884 in this case requires an independent 6.3-volt supply.

To operate the circuit, vary the voltage between the cathode and the grid-plate while monitoring the current with the ammeter. When the tube is operating properly, the grid is positive with respect to the cathode and electrons thermally emitted from the cathode are accelerated toward the grid. Along the way the electrons may strike argon atoms with which the tube is filled at low pressure. At low voltages the collisions are elastic, and the argon atoms are not altered. At any potential V the electrons have a kinetic energy

$$\frac{mv^2}{2} = Vq_e,$$

which they may transfer to the argon atoms by collision.

As the grid-plate is made increasingly positive the electrons finally acquire enough energy to ionize argon atoms with which they collide. The cathode-grid potential difference at which this occurs is called the ionization potential, V_i. Thus V_iq_e is the minimum energy sufficient to ionize the argon atoms, i.e., to remove an electron.

Experimentally you recognize this voltage by the rapid increase in the ammeter reading and by the sudden glow of light that appears in the center of the tube at the same moment.

The sudden onset of light in the argon and the sharp increase in anode current occur simultaneously as the critical potential V_i is exceeded. This is strong evidence that the argon gas is in an entirely different condition. That this is a condition of ionization seems fairly clear. It is also reasonable to assume that V_i is nearly equal to the ionization potential of the argon.

The ionization potential for argon is 15.7 volts. At this point the kinetic energy of the ionizing electrons is

$$\frac{mv^2}{2} = V_iq_e = 15.7 \times 1.60 \times 10^{-19}$$
$$= 2.51 \times 10^{-18} \text{ joules}$$

The implication of this demonstration is that electrons are bound to argon atoms by a definite binding energy which is being measured by the energy V_iq_e needed to remove them altogether from the argon atoms.

Since the Rutherford atom model has already been discussed, it is worth pointing out that the glow is caused by the emission of quanta as argon atoms recapture their lost electrons.

Students may wish to plot grid voltage (reading of voltmeter) against plate current (reading of ammeter) in order to see on a graph the abrupt increase of current.

There is a small current even at voltages far below the ionization potential, since, of course, electrons will make their way across the tube to the grid-plate so long as it is positive.

There is a similarity between this demonstration and the Franck-Hertz experiment. In that experiment the various excitation energies of mercury vapor were measured instead of the ionization potential of argon. It may be worth referring to the more detailed description of the experiment in Unit 5, Sec.

L46: Production of Sodium by Electrolysis

Solid-state rectifiers are used in a bridge circuit (Fig. 1). The dc potential difference between electrode and crucible was about 15 volts; the current was 18 amperes. In the film the rectifier circuit is seen mounted on a copper heat-dissipating plate.

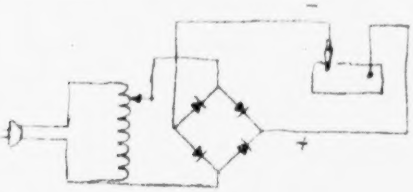


Fig. 1

One might expect water vapor in the atmosphere to react with the film of sodium on the surface of the melt. Evidently there is sufficient updraft of heated air near the surface to prevent this from happening.

L47: Thomson Model of the Atom

The horizontal component of the field of the electromagnet gave rise to an inward force on the upper poles of the floating magnets; the outward force on the lower poles was weaker since they were further from the electromagnet. Thus the electromagnet exerted a net inward force on the ping-pong balls. The model was not exact, since the net force on the magnets was only approximately proportional to the distance away from the center of the pattern.

L48: Rutherford Scattering

The law of areas holds because the nuclear repulsion is a central force. This conclusion is true whether or not the force is a Coulomb inverse-square force. (See the notes for Loop 15, Central Forces, and text Section 8.4.)

A student may observe that the alpha particle slows down as it approaches the nucleus (the displayed points are closer together). This is easily interpreted; consider the extreme case when the particle is aimed head-on at the nucleus, slows down, stops, and then moves out again. In general, for any path, the alpha particle gains potential energy in coming close to the nucleus, hence its kinetic energy must decrease.

The student may feel that the slowing of the particle conflicts with the law of areas, since in the case of planetary motion the planet speeds up when it is closest to the sun. (But the planet loses potential energy as it "falls" toward the sun!) Have the student draw a path such as Fig. 1. (It need not be a perfect hyperbola.) The long triangle ANB has altitude h. Now the law of areas can be used to find the distance A'B'; it is less than AB because the height h' is greater than h. Thus the law of areas does not conflict with the slowing down of the alpha particle as it approaches the nucleus.

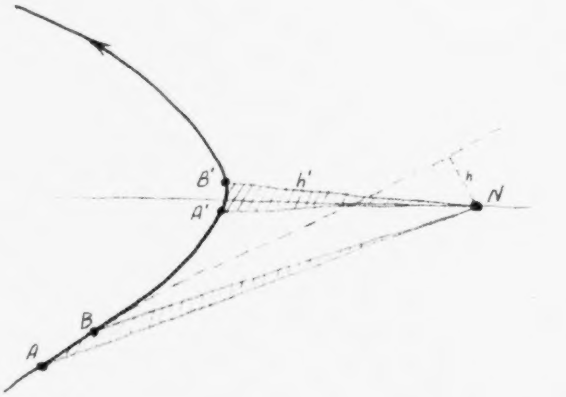


Fig. 1

You might wish to reconsider Experiment 20 with the central force now being repulsive. Students might develop various paths for different initial motions of the intruding particle.

E39: The Charge-to-Mass Ratio for an Electron

The experiment described here will give only an order of magnitude result for q_e/m (see sample data at end of this note), but students can get a lot of satisfaction out of using their own home-made beam tubes.

For this experiment it is essential to have an electron beam tube that will give a fairly well defined and clearly visible beam. Procedures are not described in great detail. The experiment is recommended only for more enterprising and resourceful students.

For other methods see:

Lecture-Experiment "Charge-to-Mass Ratio for Cathode Rays" on pages 63-66 of this Teacher Guide.

The experiment "The Mass of the Electron" in Laboratory Guide to Physics, Second Edition, Physical Sciences Study Committee.

R. M. Sutton, Demonstration Experiments in Physics (McGraw-Hill).

J. Barton Hoag, Electron and Nuclear Physics (van Nostrand).

The Taylor Manual of Laboratory Experiments in Physics (Addison-Wesley, Reading, Mass.), which describes six methods.

Any wire between #18 gauge and #28 could be used. About 40 feet are needed for each pair of coils.

After students have discussed and performed the Millikan experiment they can calculate m (the mass of an electron) by combining the results for q_e/m and q_e.

Sample Results

Accelerating voltage: 150V.

With deflecting plate also 150V above filament potential, beam goes straight up the tube.

With deflecting plate connected to ground, a current of 0.94 amps in the coils is needed to straighten out beam.

With magnetic deflection above, beam hits plate 2.5 cm from hole in a node (distance d in Fig. 5).

Distance between anode plate and deflecting plate = 1.5 cm (2x in Fig. 5).

$$\therefore R = \frac{d^2 + x^2}{2x} = \frac{6.25 + 0.5}{1.5} = 4.5 \text{ cm}$$

$$= 4.5 \times 10^{-2}\text{m}.$$

$$\text{Field } E = \frac{V}{2x} = \frac{150}{1.5 \times 10^{-2}} = 10^4 \text{ V/m}.$$

From calibration curve for coils,

$$B(I = 0.94 \text{ amps}) = 5.2 \times 10^{-4} \text{ tesla}.$$

$$\therefore q_e/m = \frac{E}{B^2R} = \frac{10^4}{(5.2 \times 10^{-4})^2 \times 10^{-2}}$$

$$= .82 \times 10^{12} \text{ coul/kg}$$

$$= 8.2 \times 10^{11} \text{ coul/kg}.$$

E40*: The Measurement of Elementary Charge

Chapters 17 and 18 of the text follow in historical sequence the development of atomic theory from the laws of chemical combination to Thomson's model of the atom as a positive blob embedded with negatively charged electrons. In Chapter 17 Faraday's work on electrolysis is presented as evidence for a connection between electricity and matter. In Chapter 18 the early work on cathode rays is followed by a description of Thomson's determination of q_e/m and Millikan's experiment to determine the value of the charge on the electron.

However, we suggest that in the lab the Millikan experiment be done before the experiment on electrolysis. Students can then use the value of q_e to determine the mass and approximate size of an individual metal atom deposited in electrolysis.

Q1 Field $E = \frac{V}{d}$ volts/meter

Q2 Electric force $F = qE$ newtons

Q3 Gravitation force $F = ma_g$ newtons

This experiment is a modification of Millikan's experiment. Its purpose is to measure very small charges and to consider answers to the questions: "Is there a limit to how small an electric charge can be?" and if so, "Does electric charge come in multiples of some basic 'atom' of electricity?"

One should ask students how the attribute of "smallest" can be demonstrated. Perhaps the answer to this question can be left for the experiment to clarify. Or perhaps it will help to illustrate the question by means of the following analogy: you have a collection of cardboard egg cartons. Each of these essentially weightless containers has concealed in it anywhere between none and a dozen eggs. How do you now find that an "egg" is not endlessly divisible, but comes in multiples of a "smallest possible" chunk? And how do you find the size (here the weight) of such a chunk? The rather obvious answers to these questions, achieved with the help of a balance, are analogous to the answers to our questions about electric charge achieved with the help of the instrument described below.

Students may ask, "How do they ever manage to make batches of latex particles in which all particles have the same size?"

It seems that "seed" particles are added to a mixture of monomer (i.e., styrene, if the polymer being made is polystyrene) and catalyst. The seed particles act as nuclei for growth of larger particles of polymer. Now, it turns out that the rate at which a particle grows depends on its size, and the smaller the particle the faster it grows. The result of this is that there is a "sharpening" of the size distribution. The "seeds" need not be all the same size, as long as they are small. Of course, there should be no new nucleation once the process has started. Soap is added to prevent the formation of new particles, and to prevent coagulation of particles already formed.

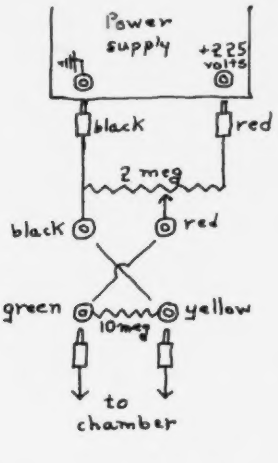
References:

Bradford, et al., J. Colloid Science, 11, 135 (1956).

Vanderhoff, et al., J. Polymer Science, 20, 225 (1956).

The student instructions assume that the apparatus has been set up and put in working order. For details of how to assemble and adjust the apparatus, refer to the note packed with the equipment and reprinted at the end of these notes.

If the instrument has not been used for some time, it is a good idea to shake the plastic bottle to make sure the suspension is well mixed. If it has lost water by evaporation, add a little distilled water.



If students have difficulty seeing the particles let them introduce a little smoke into the chamber. (You don't have to encourage cigarette smoking: draw some smoke from a just-extinguished match into a medicine dropper, and expel it into the chamber.) When most of the smoke has settled students should have no difficulty in seeing the tiny smoke particles, and can adjust the light source and microscope for maximum visibility. Then go back to latex spheres. If still none are visible the atomizer may be at fault rather than the optics.

Q4 Some appear to move up, others down in the electric field.

Q5 Some particles are positively charged, others negatively.

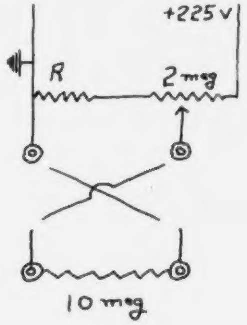
Q6 These are the more highly charged particles.

Q7 Rapidly moving particles have higher charge.

Also, the evidence that charge is quantized is less convincing if one works with highly charged particles.

The "balancing voltages" for particles carrying six and seven units of charge are 39 and 34 volts respectively. The difference is probably not significant experimentally. On the other hand the balancing voltages for singly, doubly and triply charged particles are 234, 117 and 78 volts respectively, and the difference is much clearer.

You can make it impossible for students to work at uselessly low voltages by adding a fixed resistor between potentiometer and the black input terminal.



If R is 250K the minimum voltage will be about 25 volts.

One group of students may not accumulate enough data to give clear evidence of the quantization of charge. If data from the whole class are pooled the quantization should be more obvious. But before pooling data try to make sure that there are no systematic errors in the results of any group. The data might be graphed by each lab group before pooling their results. This makes it easier to identify any systematic errors between different groups. It might be well, for example, to check the voltmeters against one another.

Some Typical Data

The results given here were obtained with spheres having a diameter of 1.305 microns. These are not currently available. The balancing voltages for the smaller spheres (1.099 micron diameter) now supplied will of course be smaller. The values of V will be reduced in the ratio $(\frac{1.1}{1.3})^3 = 0.61$, and the values of $\frac{1}{V}$ will be increased by the factor $\frac{1}{0.61} = 1.64$.

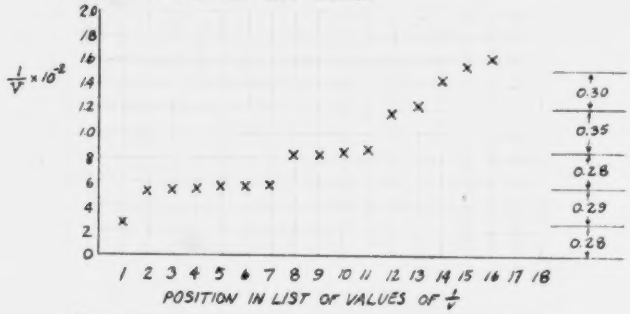
Balancing Voltages*

V	$\frac{1}{V} \times 10^{-2}$
11 volts	9.1 volt^{-1}
80	1.25
115	0.87
120	0.83
85	1.18
35	2.86
112	0.89
175	0.57
175	0.57
60	1.67
68	1.47
170	0.59
360	0.28
175	0.57
65	1.54
170	0.59
170	0.59
120	0.83

Values of $\frac{1}{V}$ Arranged in Increasing Order

$\frac{1}{V} \times 10^{-2}$	Position in list
0.28 volt^{-1}	1
0.57	2
0.57	3
0.57	4
0.59	5
0.59	6
0.59	7
0.83	8
0.83	9
0.87	10
0.89	11
1.18	12
1.25	13
1.47	14
1.54	15
1.67	16
2.86	17
9.1	18

Graph of values of $\frac{1}{V} \times 10^{-2}$ plotted against the number defining the position of each value on the list:



Remember that charge q is directly proportional to 1/V. The graph could be drawn to show values of q directly instead of 1/V, but this requires more calculations before plotting. Whichever way it is done, the graph of pooled results should show vividly that the charge on each particle is a whole number of "smallest observed" units. If we measure from the graph the size of this "smallest observed" unit, we find in our example

$$ma_g = 1.18 \times 10^{-14} \text{ nt}$$

$$d = 5 \times 10^{-3} \text{ m}$$

$$ma_g d = 5.9 \times 10^{-17} \text{ joules}$$

$$q = \frac{ma_g d}{V} = 5.9 \times 10^{-17} \times 0.28 \times 10^{-2}, \text{ so}$$

$$q = 1.6 \times 10^{-19} \text{ coulombs.}$$

This value of $q = q_e$ is considered to be the elementary unit of charge. The elementary particle carrying it is then called the electron.

Notice that even if we did not have the lowest reading of 1/V we could still have found a value for q by calculating the charge that corresponds to the difference in 1/V values between successive steps on the graph. In this example we would probably have taken 0.29×10^{-2} volt^{-1} as the constant value of the step. This would give a value of $q = 1.7 \times 10^{-14}$ coulombs for the quantum of charge.

For particles having diameter 1.099 microns the corresponding values are

$$ma_g d = 7.3 \times 10^{-15} \text{ nt},$$

$$ma_g d = 3.6 \times 10^{-17} \text{ joules},$$

and the difference in 1/V values between successive steps in the graph should be 0.63×10^{-17} volt^{-1}.

---

*Diameter of spheres = 1.305 micron

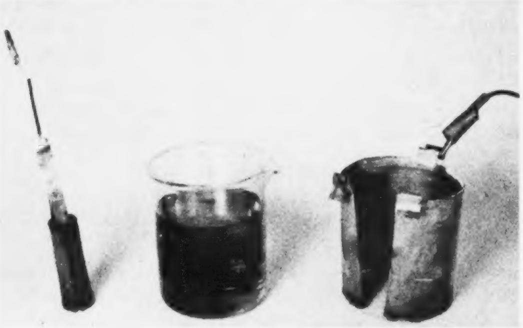
Discussion

Notice that the idea that 1.6×10^{-19} coulombs is the smallest possible charge rests only on the fact that despite an enormous number of measurements like the foregoing, no smaller charge has been observed. There is no other basis for the idea that this must be the smallest charge.

This may be a good place to point out to the class that any physical quantity that exists in "smallest possible" chunks is said to be quantized. We have seen that mass is quantized. We shall presently see that energy is quantized, too. Note also how, on the scale of everyday sizes of things, mass, charge, and energy do not appear to be quantized. Thus the everyday world observed with our unaided senses is remote from this aspect of the "real" world. As an example, in an ordinary 110-volt, 100-watt lamp bulb 6×10^{18} separate elementary charges enter and leave the bulb each second.

E41*: Electrolysis

We recommend that this experiment be done after Experiment 40, "The Measurement of Elementary Charge." Students will then know the value of the charge on the electron ($q_e = -1.6 \times 10^{-19}$ coulombs) and they can use the measurement of the mass of metal deposited by the passage of a known quantity of charge to calculate such things as the mass and volume of a single metal atom. If, in the experiments, we follow the strict historical sequence of the text (Faraday's work on electrolysis preceded Millikan's oil-drop experiment by about 80 years) students would use electrolysis to determine the faraday (F). Although this is an important quantity, its significance would emerge only after doing a series of experiments with many different elements and finding that the same quantity of electricity (F = 96,540 coulombs) will deposit, or release, the gram equivalent weight of any element.



Equipment

One 500 or 600 ml beaker

One sheet of copper sufficient to line the inside wall of a beaker (anode), with a connecting tab

One sheet of copper sufficient to form a cathode cylinder 6 to 8 cm long and about 3 cm in diameter, with a connecting tab

A balance from which the cathode can be suspended. If an equal-arm balance is used it should have a shelf to support the beaker (Fig. 1a). If a triple-beam balance is used it should be raised above the bench and the cathode suspended below the balance (Fig. 1b).

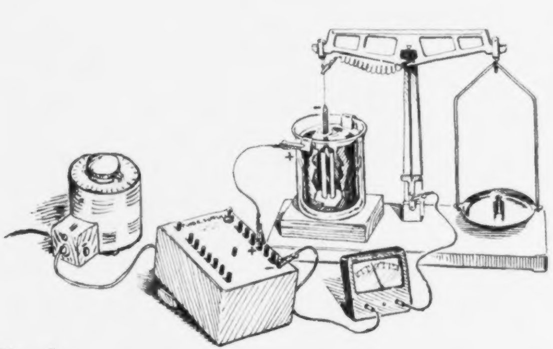


Fig. 1a

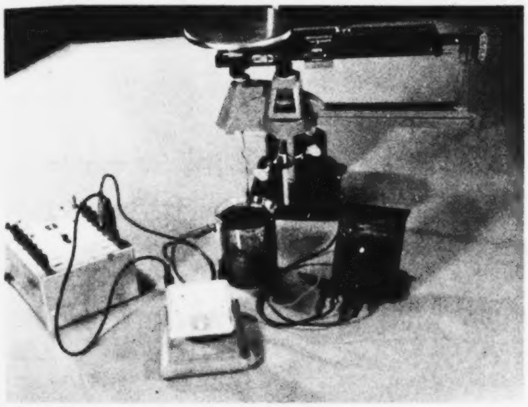


Fig. 1b

Electrolyte solution of saturated copper sulfate in distilled (or de-ionized) water with two or three drops of concentrated sulfuric acid. Both the use of distilled water and the addition of sulfuric acid are essential to the formation of a good adherent deposit.

DC ammeter 0-5 or 0-10 amperes

6-8 volt dc supply

Variable autotransformer or rheostat

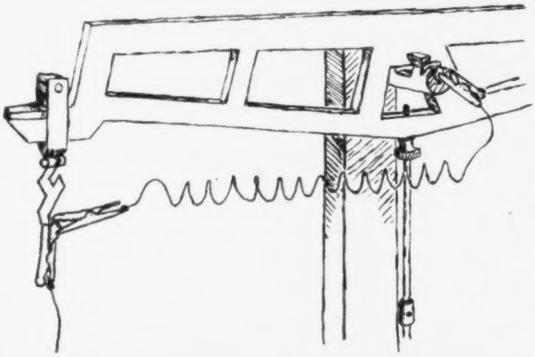
Connecting wire including a short wire connected to cathode and ending with clip to hand cathode from balance arm. Also a short wire (coiled) with clips on each end, if needed, to connect cathode to balance arm (see Fig. 2).

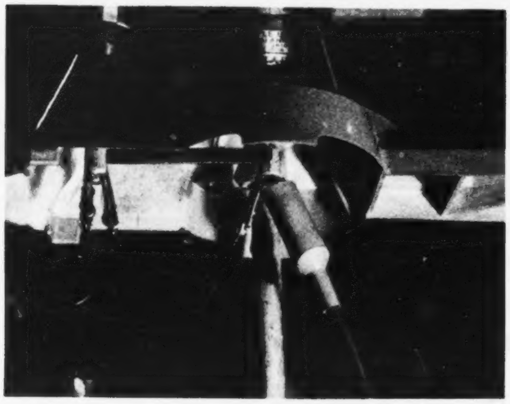
Stop watch, optional

Procedure

An unusual and very convenient feature of this experiment is that the cathode is supported in the electrolyte from the balance beam (Fig. 1). Thus the cathode need not be removed from the cell for weighing, and we have eliminated the risky step of drying it after the experiment.* Both equal-arm and triple-beam balances can be used.

Another refinement is to control the current by means of an autotransformer ("Variac" or "Powerstat"—see Fig. 1) which provides current control over a wider range than the more conventional rheostat in series with the cell. Of course a rheostat can still be used if more convenient and should be connected in series with the output of the power supply.





If an ordinary power supply is not available one can be made from a bell transformer with a rectifier in series with its output. The ammeter in the circuit will read the average of the resulting pulsating direct current, which gives correct results.

Notice that the electrical connection to the cathode must be made through the wire by which it hangs from the balance beam. The knife edge and its seat are not electrically conducting so they must be bypassed as indicated.

Even if the pivots of the balance are made of metal you cannot pass a 5 amp current through them.

The anode connection is made in any convenient manner.

Care must be taken to see that neither solder nor any battery clips touch the copper sulfate electrolyte since some foreign metal will dissolve and, by replacing copper atoms in the electrolysis process, will drastically alter the results. It is for this reason that both electrodes must have protruding tabs for electrical connections.

Since the rate of deposit of copper will not be constant if the cathode surface is dirty or impure, it is a good idea to form a deposit of pure copper on it by a preliminary run of 10 or 15 minutes. During this time the controls can also be adjusted to obtain the desired current. Explain to the students why they can start from any state of the electrode, not just the unplated electrode at t = 0.

$$\rho_c = 8.9 \text{ g/cm}^3$$

$$\rho_e = 1.3 \text{ g/cm}^3$$

$$\therefore 1 - \frac{\rho_e}{\rho_c} = 1 - \frac{1.3}{8.9} = 0.85$$

$$\therefore \quad \Delta m = \frac{\Delta m'}{0.85}$$

Q1 $Q = I \times t$. (A typical answer—a current of 5 amps for 10 minutes—would be $5 \times 10 \times 60 = 3{,}000$ coulombs. This would give a true mass increase of 1.02 g.*)

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*If you still have trouble due to some of the deposited copper dropping off the cathode, try reversing the current and have students measure the loss in weight of the anode.

*Theoretical value. In practice the mass increase is usually 3% - 8% lower than theoretical.

Q2 Two electrons:

$Cu^{++} + 2e^- \rightarrow Cu$

Q3 Number of electrons $= \dfrac{Q}{1.6 \times 10^{-19}}$

(For the example given above,

$$\frac{Q}{1.6 \times 10^{-19}} = \frac{3{,}000}{1.6 \times 10^{-19}} = 1.9 \times 10^{22}$$

electrons.)

Q4 Number of copper atoms deposited

$$= \frac{\text{number of electrons}}{2}$$

(0.95×10^{22} = this example).

Q5 Mass of each atom $= \dfrac{\text{mass deposited}}{\text{number of atoms}}$

$$\frac{1.02\ g}{0.95 \times 10^{22}} = 1.07 \times 10^{-22}\ g$$

Q6 Number of atoms in a penny

$$= \frac{\text{mass of penny}}{\text{mass of each atom}}$$

$$\frac{3}{1.07 \times 10^{-22}\ \text{atoms}^{-1}} = 2.8 \times 10^{22}\ \text{atoms}$$

Q7 Volume occupied by copper atom

$$= \frac{\text{volume of penny}}{\text{no. of atoms in penny}} = \frac{0.3\ cm^3}{2.8 \times 10^{22}}$$

$$= 1.07 \times 10^{-23}\ cm^3.$$

Q8 No. of atoms in gram atomic weight

$$= \frac{\text{gram atomic weight}}{\text{weight of atom}} = \frac{63}{1.07 \times 10^{-22}}$$

$$= 5.9 \times 10^{23}.$$

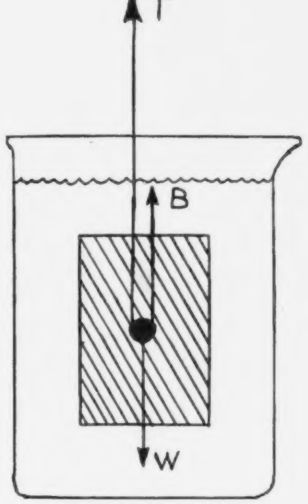
Avogadro's number is the number of atoms in one gram atomic weight.

Derivation of the Buoyancy Correction Factor

In this experiment the cathode whose mass is m and whose density is ρ_c has a volume $V = m/\rho_c$. It is submerged in a liquid (the electrolyte) whose density is ρ_e. The three forces on it are shown in Fig. 5 where T is the upward force exerted by the balance and B is the buoyant force exerted by the liquid.

When the cathode is in equilibrium $T + B = W$.

Now by Archimedes' Principle the buoyant force B is simply the weight of the liquid displaced by the volume $V = m/\rho_c$ of the cathode. And since the liquid's mass is $\rho_e V$



$$B = \rho_e V_g = \rho_e \left(\frac{m}{\rho_c}\right) a_g.$$

Also T is the apparent weight of the cathode $m'a_g$ and W is the true weight of the cathode ma_g.

Putting these values of B, T and W into the first equation above, we get

$$ma_g + \frac{\rho_e m a_g}{\rho_c} = ma_g$$

Dividing both sides by ma_g

$$\frac{m'}{m} + \frac{\rho_e}{\rho_c} = 1$$

or

$$m = \frac{m'}{1 - \rho_e/\rho_c}$$

Of course the experimenter is not interested in the total mass m of the cathode. He is interested in the increase in mass Δm, but this is proportional to the increase in apparent mass $\Delta m'$. So

$$\Delta m = \frac{\Delta m'}{1 - \rho_e/\rho_c}\ .$$

E42*: The Photoelectric Effect

Do the qualitative demonstrations (D55) using an electroscope, before students start the experiment. The demonstration shows that negative charge (electrons) can be driven off from a clean metal plate by light, and that the effect depends on the wavelength of the light and the nature of the metal surface.

One transparency (T36) can be used to explain the construction of a photocell and the measurements to be made in the experiment.

The central purpose of the student experiment is to give evidence that the wave model of light is inadequate.

Specifically the experiment shows that the kinetic energy of photoelectrons knocked out of a photosensitive surface depends on the color (and hence the frequency) but not on the intensity of the incident light. According to the wave model the kinetic energy depends on the intensity of the light.

The experiment then goes on to show that the maximum kinetic energy of the electrons is a linear function of the frequency of the incident light. A graph of energy against frequency is a straight line whose slope (as found by precise measments) is Planck's constant h. Measurements of h in this experiment are within an order of magnitude of the accepted value $h = 6.62 \times 10^{-34}$ joule·sec.

Equipment

Phototube unit—see Equipment Note on phototube unit, page 85. For this experiment the jumper leads should be uncrossed, so that as the potentiometer knob is turned clockwise the potential of the collector changes from up to a maximum of 2 volts negative with respect to the emitter.

One cannot really see the phototube inside the box, and students may not have seen one before. To prevent the experiment from becoming too mysterious, show them an unmounted phototube before they begin the experiment, and point out the emitting surface and collecting wire.

Amplifier/power supply. The case of the phototube unit is connected to the ground terminal (black) of the amplifier via the screen of the cable. In an experiment like this where small signals are amplified, noise is always likely to be a problem. In general, grounding the case of the phototube unit in this way will reduce noise. But in some instances it may be better to unground the amplifier by using a three-to-two adapter plug to connect it to the line.

Earphone or loudspeaker. At this frequency the human ear can detect ac currents as small as 1 microampere in the earphone. Since the maximum gain of the amplifier is a hundred, currents of 10^{-8} amps can be detected. The loudness of hum in the earphone or speaker increases with the current in the photoelectric cell.

Cathode ray oscilloscope can be used to detect the amplified photocurrent instead of earphone or loudspeaker. It is important to use shielded connecting wires to avoid extraneous pick-up. Set the 'scope to a sweep rate of a few hundred per second; set vertical gain to maximum.

Light source. Mercury vapor lamps are the best for this experiment (e.g, Macalaster #3400, $10.00; or the small 4-watt "ozone" lamps made by General Electric—which cannot be run without ballast). These emit the four frequencies listed. Because they emit ultraviolet, students must not look directly at them.

Fluorescent lamps give a continuous spectrum with bright mercury lines superimposed on it, as can be seen easily with a pocket spectroscope. Fluorescent room lighting is adequate if the photocell is directly below the ceiling fixture. A large lens can be used to concentrate more light on the cell.

Fluorescent desk lamps may be unsatisfactory unless you can screen out the large inductive hum they cause in the earphones which has nothing to do with the amount of light shining on the photocell.

An incandescent light source (such as the light source of the Millikan apparatus) is less satisfactory. It gives a continuous spectrum and so filters whose cut-off frequencies are exactly known must be used (see section on filters, below). If earphone, loudspeaker or oscilloscope is going to be used as detector the beam must be "chopped" to give an ac current in the phototube. The 12-slot strobe disc driven by the 300 rpm motor gives a 60-cycle per second signal which is unsatisfactory because it can easily be confused with line frequency pick-up. You can make a cardboard disc with about 60 teeth which will give a chopping rate of 300 per second when mounted on the 300 rpm motor.

An uninterrupted beam of light from an incandescent lamp—or daylight—can also be used: the photocurrent will be dc and so will the amplified current. Use a dc milliameter or 0-2.5V dc voltmeter instead

NOTE: Many of the details of these notes are appropriate only to the equipment supplied in 1967-68. New notes will accompany (or precede) the new equipment. The less detailed Student Handbook notes have been adapted to the new equipment.

of speaker or oscilloscope to detect the amplifier output. Set the dc offset control very carefully so that the meter reading is zero when the phototube is covered. This is certainly the simplest method to use and needs least specialized equipment. It has the advantage that the noise problem is largely eliminated.

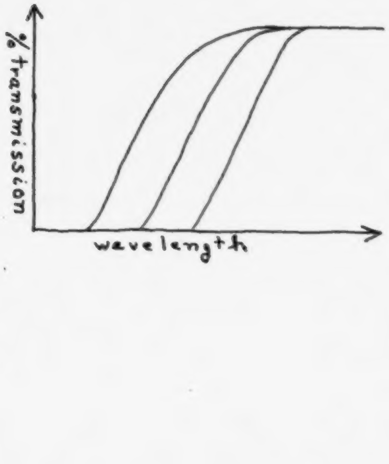
Filters to fit over the photoelectric cell window. At least three are necessary. The yellow, green and blue filters supplied by Project Physics do quite well for isolating the mercury yellow, green and blue lines. Cut the yellow filter (Wratten #22) into four pieces with a sharp pair of scissors and mount each small square in a 35 mm slide mount. With no filter, the highest frequency effective is either the violet or the ultraviolet mercury line, probably the former. If you use a fluorescent lamp source, you can assume that the dominant frequencies transmitted by the filters are the mercury lines.

Voltmeter (0-2 or 0-5 volts dc) for measuring "stopping voltages." Its use is described more fully in the procedure section below.

Students may notice that if they continue to turn the voltage control knob past the cut-off setting, the signal increases again. This is because some photoelectrons are emitted from the collector wire and, because the collector is now quite negative with respect to the emitter, they are drawn to the emitter. The current is now in the reverse direction, but this cannot be established with loudspeaker detector (though it can with oscilloscope or dc meter).

If students ask about this you can either explain the cause and tell them to find the voltage setting for minimum signal or reduce the reverse current by making a black stripe with narrow tape or grease pencil down the middle of the tube face to prevent light from falling on the collecting wire.

The Wratten series of filters have sharp cut-off at the high frequency (short wavelength) end:



By using some of these filters and an incandescent lamp (continuous spectrum) many more values of stopping voltage could be obtained. The filters are made by Kodak and can be obtained through a shop that sells photographic supplies.

Wralten Filter	Cut-off λ
Glass alone	3200A
#1 (poor stability)	3500A
2B	3900A
4 (poor stability)	4500A
8	4600A
12 + 8	5000A
15 + 8	5100A
16	5200A
23A	5600A
26	5900A
29	6100A
70	6500A
97 + 26	6800A

Q1 The stopping voltage increases as the frequency of light increases.

Q2 Yes, the greater the light intensity the louder the hum (and we can deduce that more photoelectrons are emitted).

Q3 No, stopping voltage does not depend on intensity. (But because the signal is weaker for lower light intensities it may be difficult to find precise cut-off settings at low intensities.)

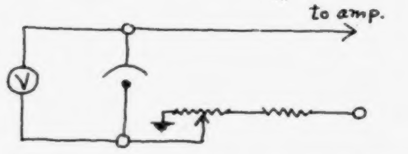
Q4 Students will not be able to detect any time delay. The wave model of light requires a delay of about a hundred seconds (see derivation on page); the particle model is consistent with instantaneous emission.

You may want to stop here. For a discussion of the meaning of the results so far, see the notes at the end of these instructions.

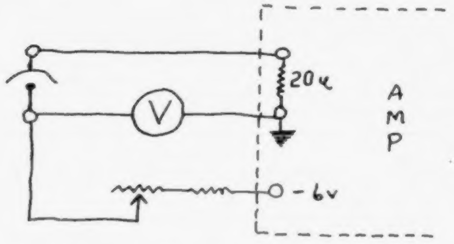
The second part of the experiment requires that students make more precise measurements of stopping potential, and then graph their data and deduce Planck's constant from it. They should have already read Sec. of the text and understood Millikan's graph on page , since this is not material that can be learned best by starting with the experiment.

Voltages measured in this way will not be absolute values. The resistance of the phototube is very high (~ 10 megohms). When a voltmeter of appreciably lower resistance is put across it the meter draws some current and voltage across the phototube drops. But the relative values for stopping voltage can still be compared and used to plot the Vq_e vs. f graph.

(Alternatively, if the voltmeter is connected between the collector of the phototube and the ground terminal of the amplifier, it will give correct absolute readings.)



a) Voltage readings are relative only.



b) Absolute voltage measurement.

$$q_e = 1.60 \times 10^{-19} \text{ coulombs.}$$

Ideally, the points should be in a straight line.

Q5 As explained in Section 18.4 of the text, the wave theory of light can neither predict nor explain the results of this experiment. In particular, if the very low intensity light strikes the photocell one would expect an appreciable time delay before the emission of a photoelectron (see page in this Guide).

Q6 The particle theory, on the other hand, can explain the results of this experiment quantitatively as well as qualitatively.

Q7 Q8 The value of h will be approximate for several reasons. For example, a fluorescent lamp gives out a continuous spectrum which contains all visible frequencies at low intensity. Some of these which pass through the filters will have frequencies greater than those of the bright emission lines. Also the electron-emitting surface is never uniformly clean. Various spots on it have various work functions. (Millikan had to prepare his pure metal surfaces by shaving off the oxide coating in a high vacuum.) The signal is small, and there may be considerable "noise" as well, which makes it difficult to make accurate determinations of stopping voltage. And the use of inexpensive voltmeters limits the precision with which you can measure the low values of stopping voltage. Results should, however, be to better than an order of magnitude. A typical plot is given at the end of these notes.

Q9 The value of f_0, the threshold frequency, and of W, the work function, varies from metal to metal and depends on the condition of their surfaces. Students' results will depend on the accuracy of their voltage readings (see note above on use of low resistance voltmeters).

Q10 No, because there is no result in this experiment that demands a wave theory for its explanation.

Discussion

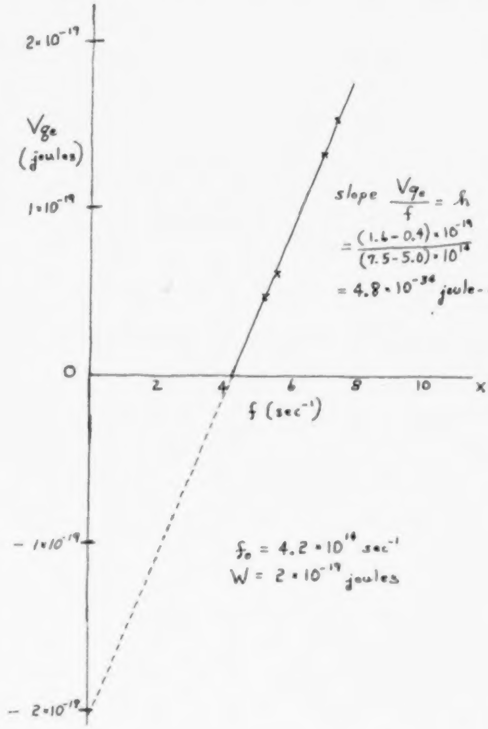
The two transparencies (T36 and T37) which show idealized results of this experiment, and the photon theory explanation can be used after students have finished the experiment.

The important ideas to emphasize in discussion of this experiment are that:

1. The stopping voltage and hence the maximum energy of photoelectrons is a linear function of the frequency, and is independent of the intensity of the incident light.

2. The photoelectrons are emitted immediately when light falls on the photoelectric cell.

3. The preceding two statements are both inconsistent with the wave model of light (text, page) according to which (1) the maximum energy of the photoelectrons should also depend on the intensity of light, and (2) the photoelectrons should not emerge until several hundred seconds **after light strikes the cell.**



$$\text{slope } \frac{Vq_e}{f} = h = \frac{(1.6-0.4)\cdot 10^{-19}}{(7.5-5.0)\cdot 10^{14}} = 4.8 \cdot 10^{-34} \text{ joule-sec.}$$

$f_0 = 4.2 \cdot 10^{14}$ sec^{-1}
$W = 2 \cdot 10^{-19}$ joules

4. The slope h of the graph line turns out to be a quantity already shown to be important by Planck's study of radiation, in which he assumed light came in discrete chunks, or quanta, of energy hf.

Two papers which give details of some of the finer points of this experiment appeared in The Physics Teacher recently: 1. A. Ahlgren: "Inexpensive Apparatus for Studying the Photoelectric Effect and Measuring Planck's Constant," The Physics Teacher, October 1963; 2. H. H. Gottlieb: "Photoelectric Effect Using a Transistorized Electrometer," The Physics Teacher, November 1965.

Photoelectric effect: sample data

color	frequency $f(\text{sec}^{-1})$	stopping voltage V (volts)	kin. energy Vg_o(joules)
yellow	5.2×10^{14}	0.30	0.48×10^{-19}
green	5.5×10^{14}	0.38	0.61×10^{-19}
blue	6.9×10^{14}	0.83	1.33×10^{-19}
violet	7.3×10^{14}	0.96	1.54×10^{-19}

E43: Spectroscopy

There are several parts to this sequence of observations.

Certainly all students should look at as many spectra as possible—bright-line, absorption, and continuous—including (a photograph of) the Balmer spectrum of hydrogen.

The work becomes more quantitative as it goes along. Some teachers may not want to pursue it to the end, which requires that students calculate wavelengths, and from them perhaps Rydberg's constant and some of the energy levels of an excited hydrogen atom.

Equipment

Replica gratings and/or "take-home" pocket spectroscopes (Damon 9071, Macalaster 2412).

Sources of line and continuous spectra such as the following:

(a) incandescent lamp (see Demonstration 56: Blackbody Radiation)

(b) spectrum tubes of various gases and power supply.

(c) flames, including Bunsen burner with various metallic salts added.

(d) concentrated solutions of colored salts, dyes, chlorophyl, etc., to show absorption.

(e) fluorescent lamp.

(f) Balmer tube (atomic hydrogen) with power supply. If you use the Macalaster MSC1300 high voltage source here be sure to remove the 6.8 megohm resistor taped to its output. Macalaster #1350 spectrum tube power supply needs no alteration.

*Polaroid camera (experimental model 002 or model 95, 150 or 800) and film (speed 3000).

Procedure—qualitative

Have the class observe as many different spectra as possible to establish (a) the existence of bright-line, absorption and continuous spectra, and (b) the qualitative observation that the red light (longer wavelength) is "bent" more than the violet by a grating. In explaining spectra it is a good idea to contrast emission and absorption spectra first of all and then contrast line and continuous spectra.

Let the students use the pocket spectroscopes outside the classroom to look at illuminated signs, street lights, sky light, moonlight, etc. The Fraunhofer absorption lines of the solar spectrum (described on text page 67) can probably be seen with the pocket spectroscopes only by looking directly at the sun through a dark filter and with a somewhat narrowed slit. Razor blades taped over the slit can narrow it very well. Remind students of the danger of looking at the sun directly.

Ask whether anyone can explain why a fluorescent lamp gives both a continuous spectrum (from coating on walls of tube) and the bright-line spectrum of mercury (from mercury vapor in the tube). This is a good point to observe that continuous spectra are emitted by solids (e.g., lamp filaments) or highly compressed gases (body of the sun), while bright-line spectra are emitted by excited gases (e.g., discharge tubes or salt sprinkled into Bunsen flame). Absorption spectra are formed when light having a continuous spectrum passes through relatively cool gases or liquids or transparent solids.

Ideally students should see a hydrogen spectrum against a black background in a darkened room. Probably only three lines (red, blue-green, violet) will be visible, in which case the idea of a regular series (Balmer's) will not be very convincing. For this reason the second part of the experiment is devoted to photography which will reveal several additional lines in the ultra-violet to which the eye is insensitive.

Discuss these observations to bring out the point that because the emission of bright-lines is evidently an atomic process, an explanation of spectra can reveal a great deal about the structure of atoms. Note in this connection how each element has its own characteristic bright-line spectrum. Note how the simplest atom, hydrogen, seems to produce the simplest spectrum.

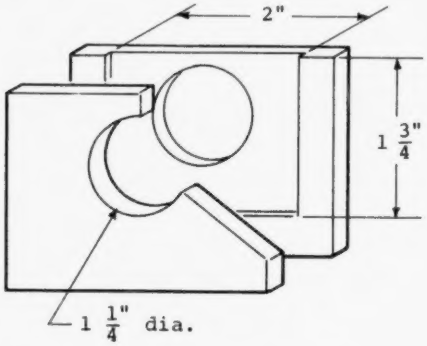
Procedure—quantitative

The hydrogen tube should give a bright red light. Old tubes give bluish light which does not produce a good Balmer spectrum.

Mount the hydrogen spectrum tube vertically against a black background in a darkened room (Fig. 1). Set a meter stick horizontally just behind the tube. Secure a grating directly in front of the camera lens, making sure that the grating lines are parallel to the spectrum tube. (You see a spectrum running horizontally.)

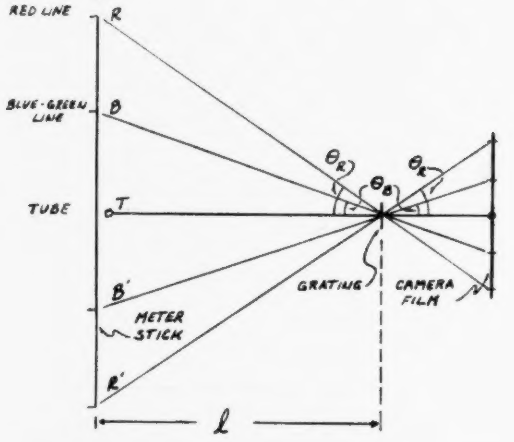
Set up the camera on a tripod or other firm support about 1.3 m in front of the tube.

An easy way to attach the diffraction grating to the lens is to use a cardboard holder:



Hole to fit camera lens (1-7/8" diameter for model 002, 1-1/4" diameter for models 95, 150, 800).

This is made with two layers of corrugated cardboard—one with a hole which fits tightly over the lens, the second with a 2" × 1 3/4" slot cut in one half of the corrugation and a 3/4" hole cut in the other half. When the two layers are cemented together, the grating is inserted and the holder can be easily positioned on and removed from the lens.



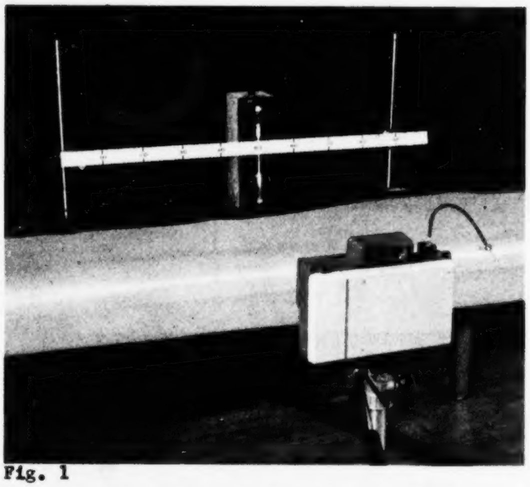


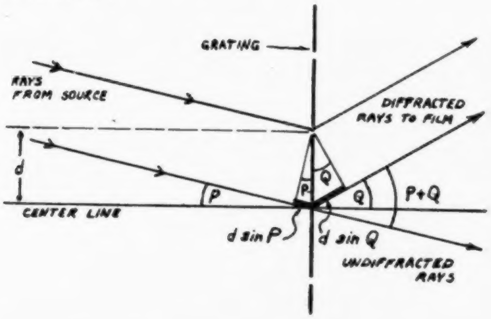
Fig. 1

The grating must of course be oriented so that its lines are parallel to the spectrum tube, i.e., vertical.

Exact exposure time will depend on the lighting conditions in your room, and is best found by experiment. You may be able to record both the spectrum lines and the meter stick scale with a single exposure in a darkened, but not dark, room (experimental model 002 camera set to "75 speed," about 3 seconds exposure). Or you may have to make a double exposure, one of the spectrum when the tube is on, the second exposure with the tube turned off and the room lights on in order to show the meter stick scale. It is not necessary to remove the grating for the second exposure.

The resulting picture should show the spectrum lines clearly against the meter stick scale.

If you have the spectrum tube directly in front of the camera lens so that the central (undiffracted) image is in the center of the picture (Fig. 1) there may not be enough space on each side of it to show the red line (longest wavelength), which is at the greatest angular distance from the source. This will certainly occur if you use the commonly available gratings with 13,400 lines/inch and any of the cameras suggested here. To get a photograph showing the H_α line, a different set-up will be required, as shown in Fig. 2. This geometry gives a picture with the source near one edge of the photograph and the first-order red line near the other. Compare the camera's field of view with the spectrum seen by the eye through the same grating to find the best orientation for the camera, or use the focussing screen to make sure that the red line will be recorded.



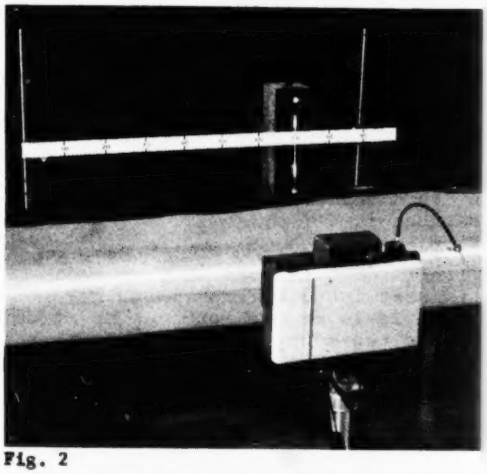


Fig. 2

Qualitative observation of the hydrogen spectrum—with or without a photograph—is as far as some teachers will wish to proceed. The important point to be made is the regularity of the spectrum, and it may be better for some students just to look.

To go further requires the presentation of the grating equation, $\lambda = d \sin \theta$, and probably a correction to it to account for the off-axis geometry of the set-up in Fig. 2.

The derivation of the grating formula $\lambda = d \sin \theta$ for normally incident parallel light is probably familiar to you, but for completeness it is reprinted at the end of this note.

The grating space d is presumably known. For the commonly available gratings of 13,400 lines/inch, $d = 1.89 \times 10^{-6}$ m.

The angle θ for each spectrum line is found from the source-grating distance ℓ and its position on the film in accordance with one or the other of the diagrams in Figs. 1 and 2.

If the source is in the center of the photograph, (i.e., light from source is normally incident on grating), the geometry of Fig. 1 applies.

It can be seen from the figure that for the red line

$$\tan \theta = \frac{RT}{\ell} = \frac{R'T}{\ell}$$

and similarly for the other lines. With more accuracy

$$2 \tan \theta = \frac{RR'}{\ell}, \text{ etc.}$$

Since RR' is identified from the photograph and ℓ (the source-to-grating distance) is known, the values of $\tan \theta$ can be found for each line.

Tables then give values of θ and also $\sin \theta$ so that values of $\lambda = d \sin \theta$ can now be found.

If the source is not at the center of the photograph, refer to Fig. 2. In this arrangement the beam from the source is no longer incident perpendicular to the grating. Hence the rays are not diffracted through an angle θ, but through an angle P + Q as shown in the figure.

The path difference between the two rays shown in the figure is:

$$d \sin P + d \sin Q.$$

The angle P is the same for all spectrum lines. It is the angle between the source S and the center of the field of view C (Fig. 3).

Angle P is easily found since its tangent is the distance SC (from the photograph) divided by the distance ℓ from source to grating:

$$\tan P = \frac{SC}{\ell}.$$

The angle Q is the angle between a

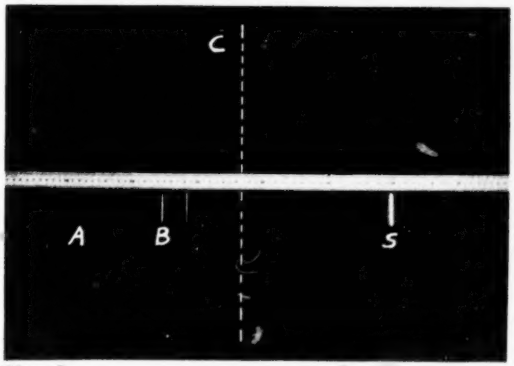


Fig. 3

given spectral line (A, B, ...) on the film and the center of the field of view C. Thus

$$\tan Q_A = \frac{AC}{\ell}, \ \tan Q_B = \frac{BC}{\ell}, \text{ etc.}$$

The procedure for finding wavelengths is as follows. From the measurements taken from the apparatus and the photograph, find tan P and tan Q for a given spectrum line. From tables find sin P and sin Q. Then

$$\lambda = d \sin P + d \sin Q,$$

or

$$\lambda = d\,(\sin P + \sin Q).$$

Students should be able to get results that agree within a few tens of angstroms (see sample data).

$R_H = 109{,}678$/cm in a vacuum. The value in air differs from this in the last two digits. This will not show up in students' results, of course.

$$E = hf = h\frac{c}{\lambda} = 6.6 \times 10^{-34} \times \frac{3.0 \times 10^{8}}{0.66 \times 10^{-6}}$$
$$= 3.0 \times 10^{-19} \text{ joules.}$$

Discussion of the hydrogen spectrum

Class discussion should bring out the point that the Balmer series is only one of several spectral series of hydrogen and that the lowest energy state of this series is not the lowest possible ("ground") state of the atom. It is next to the lowest state. When electrons do fall to the ground state, the quanta emitted have higher energy and lie in the ultraviolet (Lyman series).

Discussion should also make it clear that spectral series of other elements are generally much more complex, but from them it is also possible to work out the related energy level differences of excited valence electrons.

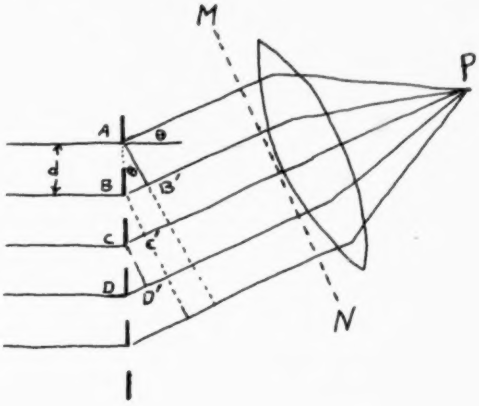
It is probably also worth reminding students that the lines of the Balmer series involve excited atoms, but not ionized atoms. If possible show pictures of the Balmer series in absorption in the spectra of stars. From this kind of evidence we know that hydrogen is by far (~90%) the most prevalent element in the universe as a whole.

Sample Results

As many as nine lines of the hydrogen Balmer spectrum have been measured by this technique. More typically, students will be able to measure about five lines:

measured value (Å)	accepted value (Å)	% difference
6350	6562	3.2
4830	4861	0.6
4430	4340	2.1
4050	4101	1.3
3920	3970	1.3
3820	3835	0.4

DERIVATION OF THE GRATING FORMUAL $\lambda = d \sin \theta$



Consider a parallel beam, normal incidence at the grating.

The diffracted rays are brought to a focus by the (eye or camera) lens. There will be reinforcement at the point P if the diffracted rays are in phase in the plane MN. For this to happen the path difference between rays diffracted at successive openings in the grating must be a whole number of wavelengths. Thus BB' = nλ; CC' = nλ; DD' = nλ, etc. for P to be bright.

If d is the grating spacing, BB' = CC' = DD' = ... = d sin θ. So point P will be bright if

$$n\lambda = d \sin \theta.$$

There are a series of values of sin θ corresponding to n = 1, 2, 3... which will satisfy this equation for a given λ. In this experiment we are only concerned with the first order spectrum (n = 1) for which the formula simplifies to

$$\lambda = d \sin \theta.$$

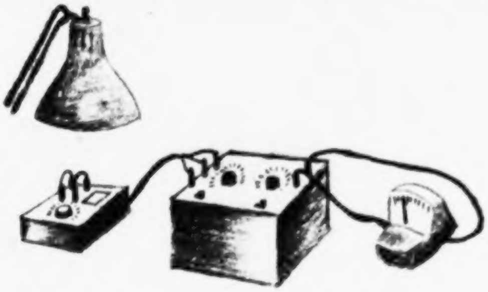
Longer wavelengths are diffracted through larger angles ($\sin \theta \propto \lambda$).

E42* Addendum

The Photoelectric Effect Made Simple

You may be bewildered by the number of alternative light sources and detectors that are suggested for E42*, The Photoelectric Effect, in the Unit 5 Teacher Guide. The simplest setup, and one that requires the least special equipment use:

a) a desk lamp, or other incandescent source
b) voltmeter (2.5V dc) or milliameter to detect amplifier output



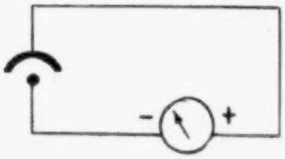
There is no noise problem with this arrangement.

Disadvantages are:

1. The incandescent lamp has a continuous spectrum which makes it difficult to know what the highest effective frequency passed by the filter is.

2. The dc offset control must be carefully set to make sure that the meter reads zero when there is no light falling on the phototube.

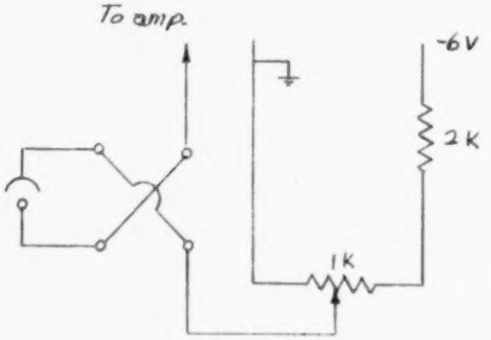
You can show the direction of the photocurrent and its dependence on light intensity very simply with the Project Physics phototube unit and microammeter, and a desk lamp. Connect the meter diretly across the phototube, using the black jacks on the front of the unit.



A 100-watt tungsten bulb about 10 cm from the phototube gave about 5 microamps.

Equipment note on the Phototube Unit (PUB-100)

The unit consists of a 1P39 gas-filled phototube, 1000-ohm potentiometer and 2000-ohm fixed resistance mounted in a metal box.



The unit is connected by shielded cable to the Project Physics amplifier/power supply unit. The box is connected via the cable screen to the ground terminal of the amplifier. (In some cases ungrounding the amplifier unit, by using a three-to-two plug to connect it to the line, may help reduce noise.)

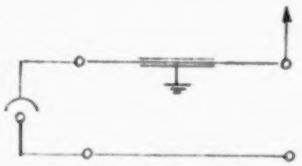
One end of the potentiometer is connected to ground, the other is 2 volts below ground. The voltage across the phototube can thus be varied between 0 and 2 volts.

For normal phototube applications (where one is interested in as large a photo current as possible) the emitter is made negative with respect to the collector; cross the jumper leads so that as the potentiometer knob is turned clockwise the potential of the collector increases up to a maximum of 2 volts above the emitter. Use the unit this way—or connect directly to the 1P39 via the two black jacks—when using it as a light sensor, for time-of-flight measurements, etc.

But in Experiment 42*, The Photoelectric Effect, a counter potential is applied across the tube (collector negative with respect to emitter), and gradually increased until the flow of photoelectrons is stopped. The jumper leads are uncrossed.

In either application the photocurrent can be increased 100X by the amplifier and detected by earphone or oscilloscope (for intermittent light sources) or meter (for steady light source).

There will probably always be some background noise due to pickup from nearby 60-cycle circuits when the amplifier is set to a high gain. When determining the stopping voltages in Experiment 42* it is important to distinguish between signal and noise. Cover the phototube window with an opaque shield. Any remaining amplifier output is noise. When doing the experiment increase the counter potential until the amplifier output is reduced to this level.



Shielding one of the jumper leads and grounding the shield to the box will help reduce pick up.

NOTE: Many of the details of these notes are appropriate only to the equipment supplied in 1967-68. New notes will accompany (or precede) the new equipment.

Millikan Apparatus

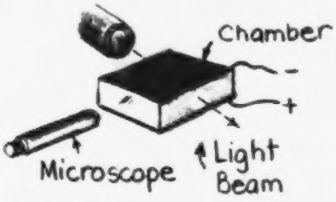
1. Connect the wire leads from the light source to a 6.3-volt, 2.5-ampere ac or dc terminal of a power supply. Turn the power supply on at once to dry out the tube. If moisture collects on the lens, remove it, dry with a soft tissue, and leave lens off until tube warms up.

2. Mount the light source on the pivot arm, in the U-shaped support. Set the pivot arm so that it makes an angle of about 50° with the instrument axis, with the light source to the left, when viewed from the front of the instrument. Adjust the source to produce a sharp vertical image of the filament on a white card held at the center of rotation of the pivot arm. Hold it in place with the aluminum clip and rubber band attached to the pins in the pivot arm.

3. Remove the small shiny reflector from the objective end of the microscope. Keep the reflector to use with the microscope to examine opaque objects. (It illuminates the object and ensures the correct distance between objective lens and object—when the end of the reflector just touches the object.)

4. Mount the microscope in the spring clip and focus it on the white card at the center of rotation of the pivot arm. The parallel lines of the scale in the microscope should be horizontal.

5. Mount the plexiglass chamber over the center of rotation of the pivot arm, fitting the locating pins into the small holes and attach with a rubber band. The hole should be to the left with the two wire leads coming from the back of the chamber. Be sure chamber walls are clean. The light beam should enter the chamber close to the corner. Connect the two wire leads to the jacks on the control box, yellow to yellow and green to green.



6. Connect leads from the control box to the dc power supply, red plug to 224V terminal, black plug to ground. CAUTION: TURN OFF POWER SUPPLY BEFORE MAKING THE CONNECTIONS. Turn on power supply. When control-box switch is in the center, no voltage is connected to the plates in the chamber. When switch arm is down, upper plate is positive; when it is raised, lower plate is positive. The knob on the right side of the control box governs the magnitude of the voltage between the plates. When it is fully counterclockwise the voltage is zero; fully clockwise, the voltage between plates is equal to the voltage of power supply.

DEMONSTRATION: Brownian Motion

The apparatus is now assembled for demonstration of Brownian motion (Unit 3). Draw a little smoke from a just-extinguished match into a medicine dropper, and expel into chamber through the hole in wall. Too much smoke will make the whole field of view look grey. If this happens, wait until most of the smoke has settled, or blow some out. You should see a relatively small number of bright spots against an almost black background. The "jiggle" of these tiny particles is very noticeable. You may want to prepare the ground a little for the Millikan experiment by showing the effect of an electric field on the smoke particles (which are electrically charged). Use the reversing switch and potentiometer to demonstrate the effect of changing field strength and direction.

7. Remove the atomizer cap (nipple) from the plastic bottle. Pour about 1 cm depth of latex particles into the bottle. It is not necessary to dilute the suspension. Push the short length of narrow white plastic tubing into the inside of the atomizer cap, and then replace the cap, with the plastic tubing inside the bottle. Attach one end of the amber latex tubing to the nipple and insert the inflating needle in the other end of the tubing. Mount the bottle in the spring clip and insert the needle into the hole nearer the microscope in the chamber wall until the push-nut touches the chamber wall.

8. Connect a high impedance dc voltmeter (at least 10,000 ohms per volt or 100 microamp full-scale deflection) to the black and red jacks on the control box. Red to +, black to -.

9. Introduce a cloud of spheres into the chamber of the Millikan apparatus by giving the plastic bottle a sharp squeeze.

10. You should see several tiny bright spots of light against a dark grey background. It may be necessary to adjust the microscope very slightly to focus on the spheres, or to adjust the light source. When the apparatus is correctly set up the beam of light enters close to the corner of the chamber, and the end of the microscope is only a few millimeters from the front wall of the chamber. The field of view will be dark grey. If you do not see any spheres after making a slight adjustment of light source and microscope, introduce a little smoke into the chamber (as in the Brownian motion demonstration). Adjust light and microscope for maximum visibility of the smoke particles. Then try again with the latex spheres. The light beam should pass between the two plates. You may find that some of the beam passes above or below one of the plates, leaving the working space not fully illuminated. If so, adjust the beam up or down slightly by placing a thin card between the light source tube and the U-shaped support.

EXPERIMENT: Measurement of Elementary charge

Details of the Millikan experiment itself are given in the Student Handbook and Teacher Guide. In the method described the electric field is adjusted until the electric force acting upwards on a particle is just equal to the gravitational force acting downwards. The same apparatus can also, of course, be used for the more conventional version of the experiment in which the two forces do not balance, and the sphere moves with terminal velocity.

Comments on the determination of relative Atomic Masses

There are three approaches that can be taken to determine relative atomic masses.

A. Chemical reaction method. The relative masses of elements entering into a reaction can be determined. In this way, for example, it can be demonstrated that in the formation of water the mass of oxygen is 7.94 times the mass of hydrogen. If one now assumes, as did Dalton, that water consists of one oxygen atom and one hydrogen atom, then one can conclude that the mass of one oxygen atom is 7.94 times the mass of one hydrogen atom. (See Sec. 17.2.)

B. Electrolysis method. A second line of evidence comes from electrolysis experiments. The decomposition of water by electrolysis produces 1.008 grams of hydrogen and 8.00 grams of oxygen. Again, the mass of oxygen is 7.94 times the mass of hydrogen. (See Sec. 17.8.)

C. Gas density method. In the discussion of the kinetic theory of gases in Chapter 12, we saw a result of a calculation by Loschmidt. He calculated the number of molecules (N) in a cubic meter of gas at 0° C. and normal atmospheric pressure. The result is

$$N = 2.687 \times 10^{25}.$$

We can utilize this number in conjunction with gas density measurements to determine relative atomic masses. The density of hydrogen is:

$$\frac{N_H \times \text{mass of one hydrogen gas particle}}{1\ \text{m}^3}.$$

Likewise, the density of oxygen under the same conditions is:

$$\frac{N_O \times \text{mass of one oxygen gas particle}}{1\ \text{m}^3}.$$

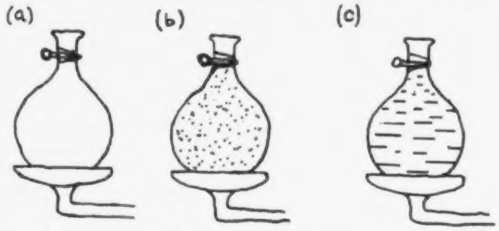
But, by Loschmidt's work, $N_H = N_O$, so the ratio of densities is:

$$\frac{\text{density of oxygen}}{\text{density of hydrogen}} =$$

$$= \frac{N_O \times \text{mass of one oxygen gas particle}}{N_H \times \text{mass of one hydrogen particle}}$$

$$= \frac{\text{mass of one oxygen gas particle}}{\text{mass of one hydrogen gas particle}}.$$

Hence, a knowledge of densities can provide relative atomic masses. Gas densities can be measured by weighing a known volume of gas. This can be done in a few simple steps. First an evacuated bulb is weighed. It is then filled with gas under measured conditions of temperature and pressure and weighed again. The difference in weight gives the mass of the gas. The bulb is then filled with water and weighed again. The difference between this mass and that of the evacuated bulb is the mass of the water. Since the density of water is known, the volume of the water, which is the same as the volume of the bulb, is determined. This volume is also the volume of the gas. The ratio of the mass of the gas to the volume of the bulb (or gas) gives the density of the gas. The densities of a number of gaseous elements at room temperature are listed in the table below.



$$\text{The ratio } \frac{\text{density of oxygen}}{\text{density of hydrogen}} =$$

$$\frac{1309}{82.5} = 15.87.$$

Therefore, from the gas density data, it is concluded that the mass of one oxygen particle is 15.87 times the mass of one hydrogen particle. If we compare this value with the value obtained by the chemical reaction method and the electrolysis method we see that our new value is twice as large. Two explanations could account for this discrepancy. Either a gas particle of oxygen contains twice as many atoms as a gas particle of hydrogen, or water contains twice as many hydrogen as oxygen atoms. In the first case the relative atomic masses of oxygen to hydrogen would be 8:1, in the second case 16:1.

If we calculate the other relative densities (density of element/density of hydrogen), we obtain the values in the last column of Table 1. Notice that the relative density for chlorine is identical with the number for chlorine listed in Table 17.4 of the text.

We now have two sets of numbers for some elements: one set obtained from electrolysis and chemical reaction measurements, the other from gas densities. To establish the relative masses of atoms, we must have some reliable way of assigning formulas to elements and compounds. Such a way was proposed in 1858 by S. Cannizzaro, who called attention to a regular pattern developed from data like those shown in Table 17.6. The second column shows values of gas densities measured at 100° C and 1 atmosphere pressure. In the third column are values of the fraction of the weight of each substance due to the presence in it of hydrogen. This fraction can be determined by measuring the amount of water obtained when each of the indicated substances reacts with oxygen. On multiplying together the values for a given gas in the second and third columns, we get the corresponding value in the fourth column. These values represent the numbers of grams of hydrogen per liter of each compound. That is:

$$\text{density}\left[\frac{\text{g compound}}{\text{liter}}\right]\times\begin{matrix}\text{weight}\\\text{fraction}\\\text{hydrogen}\end{matrix}\left[\frac{\text{g hydrogen}}{\text{g compound}}\right]$$

$$=\frac{\text{grams hydrogen}}{\text{liter}}\ .$$

In the fifth column we show that the weights of hydrogen per liter of the various compounds stand to each other in the ratio of small whole numbers. What interpretation shall we place upon this strikingly simple relationship?

Loschmidt's number is independent of the gas under investigation. It implies that under specified conditions of pressure and temperature, a cubic meter (or any volume) contains the same number of molecules. It then follows that the masses of hydrogen listed in the fourth column of Table 2 calculated for equal volumes of the various substances, also represent the masses of hydrogen contained in equal numbers of molecules of the various substances. The atomic-molecular theory then clarifies why these masses should stand in the ratio of small whole numbers. A given hydrogen-containing substance may contain one atom of hydrogen per molecule—or 2 or 3 or 4 or more atoms of hydrogen per molecule. If we denote by y the number of molecules of substance present, the numbers of atoms of hydrogen present in a unit volume of each substance must be 1y, 2y, 3y, 4y, etc. But, in that case, the numbers of grams of hydrogen present must stand in the corresponding small-whole-number-ratio 1:2:3:4:...—which is precisely the relationship expressed in the fifth column of our table. This explanation suggests the partial formulas given in the sixth column of Table 2.

This argument leads us to the somewhat surprising conclusion (see Table 2) that the gaseous particle of hydrogen is not a single atom, but, rather, a pair of atoms joined in an H_2 molecule. But, we are still unable to assign the complete formula of any compound. There is, however, no real reason for being discouraged; we can easily reach our goal by preparing a series of tables like that given above for hydrogen. Table 3 is such a table for oxygen.

As before, all entries in the fourth column are integral multiples of a minimum value (in this case, 0.52). As before, the gaseous element itself is seen to consist not of individual atoms, but, rather, of diatomic molecules (O_2). We can assign partial formulas as indicated in the last column of the table.

Putting together the findings in both our tables, we see that we have obtained the complete formula for water: H_2O. Once we have established such a formula, accurate atomic masses can be obtained from measurements of chemical combining masses. These are obtained more easily than accurate values of gas densities. As noted earlier (equation on page 12 of text), 1 gram of water contains

0.1119 gram hydrogen
0.8881 gram oxygen.

Given that the formula of water is H_2O, we see that 0.8881 g is the mass of oxygen containing only half as many atoms of oxygen as there are atoms of hydrogen in 0.1119 gram of that element. The masses containing equal numbers of the respective atoms would be 2 × 0.8881 gram oxygen and 0.1119 gram hydrogen, and the relative masses of individual atoms of the two species is given by the ratio of these two numbers: $(2 \times 0.8881)/0.1119 = 1.7762/0.1119 = 15.87/1.000 = 16.00/1.008$.

We have studied two tabulations, and have obtained the ratio of two atomic masses. Our procedure can be generalized to other elements. The elements themselves need not be gaseous (as are hydrogen and oxygen); all that is necessary is that the element in question forms a considerable number of compounds that are either gaseous or readily volatile. In that case we can, with confidence, assign to the element the minimum mass that corresponds to unit volume of a substance containing in its gaseous particle just one atom of the element in question. Given this minimum mass, and/or the assignments of partial formulas it makes possible, we can then proceed, by the three methods noted above, to assign the atomic mass of the element. In this way atomic masses can be assigned to a substantial fraction of the known elements and there are additional procedures, which we need not discuss, that make it possible to find the atomic masses of all the known elements.

Table 1: Gas Densities of Some Elements at 298° K and 760 mm Pressure

	Density g/m^3	Density Relative to hydrogen density = 1
Hydrogen	82.5	1
Nitrogen	1146	13.89
Oxygen	1309	15.87
Chlorine	2900	35.2

Table 2: Cannizzarro Method for Arriving at Formulas of Hydrogen Compounds

Substance	Gas density in gms/liter, at 1 atmosphere and 100° C	Fraction by weight of hydrogen in substance	Grams of hydrogen in 1 liter of compound	Values (in preceding column) expressed as multiples of 0.033	Partial formulas
Hydrogen chloride	1.19	0.0276	0.033	1	$H_1Cl_?$
Chloroform	3.90	0.00844	0.033	1	$C_?H_1Cl_?$
Hydrogen	0.066	1.00	0.066	2	H_2
Water vapor	0.59	0.112	0.066	2	$H_2O_?$
Ammonia	0.556	0.178	0.099	3	$N_?H_3$
Methane	0.525	0.251	0.132	4	$C_?H_4$
Ethyl chloride	2.11	0.0781	0.165	5	$C_?H_5Cl_?$
Ethane	0.982	0.201	0.197	6	$C_?H_6$

Table 3: The Formulas of Some Oxygen Compounds

Substance	Gas density in gms/liter, at 1 atmosphere and 100° C	Fraction by weight of oxygen in substance	Grams of oxygen in 1 liter of compound	Multiples of 0.52 gm per liter	Partial formulas
Nitric oxide	0.98	0.533	0.52	1	$N_?O_1$
Water vapor	0.59	0.888	0.52	1	$H_?O_1$
Carbon monoxide	0.915	0.571	0.52	1	$C_?O_1$
Carbon dioxide	1.44	0.727	1.04	2	$C_?O_2$
Oxygen	1.04	1.00	1.04	2	O_2
Sulfur trioxide	2.61	0.600	1.57	3	$S_?O_3$

Articles
Spectroscopy
Rutherford Scattering

Spectroscopy

A. Background and Scope of the Technique

Almost everyone has seen at least one spectrum, a rainbow. The band of colors, arranged from red to violet, is the visible portion of the sun's spectrum. The sun's spectrum can also be viewed by means of a spectroscope. In 1666, when he was 23, Sir Isaac Newton constructed the first man-made spectroscope.

Spectroscopy is that branch of physics and chemistry that studies the absorption and emission of electromagnetic radiation by matter. Radiation of all wavelengths can be used. Matter in all physical states, gaseous, liquid and solid, can be studied. A range of information can be obtained depending upon the wavelength of electromagnetic radiation used. Table 19.1 in text, page 69, gives a sample of information obtainable from spectral analysis.

B. Instrumentation

The basic instrument consists of three parts: a radiation source, a dispersive device and a detector.

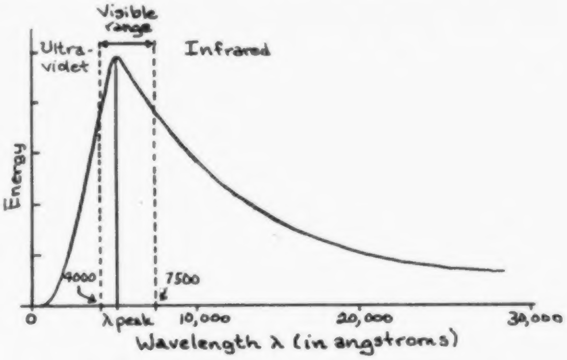
1. Radiation source: The source employed depends upon the problem being studied. There is no universal source. For example, gaseous discharges or arcs give rise to visible and ultraviolet radiation, tungsten lamps to visible, glowing filaments to infrared, and klystrons to microwave radiation. A wide variety of sources is used for different purposes.

2. Dispersive device: Two principal dispersive devices are used: a prism and a diffraction grating. Each of these will disperse, or spread out, radiation into a spectrum, thus giving intensity information as a function of wavelength. The difference between an expensive and cheap instrument is often the quality of the dispersive device. The more detail one wishes to resolve, the better the dispersive device should be. Thus, for example, to resolve the sodium doublet (see page 67 of text), the lines of which differ in wavelength by only 5.970 Å, a good instrument is required.

3. Detector: The type of detector used depends upon the radiation source being used. For work in the ultraviolet and visible regions, photographic film is useful. Thermocouples are used to detect infrared radiation, while crystals are used in the detection of microwaves. In the latter two cases, electrical signals produced at the thermocouple or crystal must be amplified electronically. Often these amplified signals are then recorded on a stripchart recorder or viewed on an oscilloscope.

C. Spectra Types

1. Continuous spectra: Continuous spectra include all wavelengths between certain limits, e.g. a rainbow. A continuous spectrum results from heating a solid, liquid, or gas under high pressure (like the sun) until it glows brightly. Such a glowing object makes a good source of radiation for spectroscopic work. In the figure below, it is seen that an object heated to 6000°K radiates with a maximum intensity in the visible region. (The tungsten filament of a photo-flood lamp is approximately 6000°K when in use.)



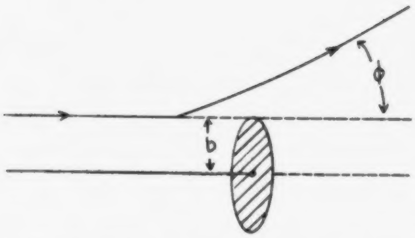
2. Discrete spectra: These are spectra in which only certain wavelengths appear. Line spectra and band spectra are discrete spectra.

   a. absorption spectra—When light is passed through a medium that absorbs radiation of specific wavelengths, the spectrum of the transmitted light is known as an absorption spectrum.

   b. emission spectra—These spectra are formed when gases under low pressure radiate.

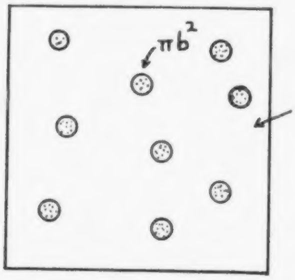
Rutherford Scattering

From Fig. 19.6 of the text it is seen that the larger scattering angle results when the alpha particle is incident along a line close to the nucleus. Looking at the figure below, we can say that any alpha particle whose undeflected path touches the circumference of the circle will be scattered through an angle ϕ. Furthermore, any alpha particle whose undeflected path intersects the circle (of radius b) itself will be deflected by an angle greater than ϕ. Thus, the alpha particle must strike the area πb^2 to be scattered by an angle greater than ϕ.



Now we can ask, what is the probability that an alpha particle will be scattered by an angle greater than φ? Each nucleus presents a target of size πb^2. The probability of an alpha particle hitting one of the targets is proportional to the total target area—that is, the total shaded area in the figure below. If there are n nuclei per unit volume, then the total shaded area is $n\pi b^2 tA$, where t is the thickness of the foil and A is the total foil area. We are assuming that the thickness of the foil is small enough that target areas do not overlap; in other words, we are considering that only single scatterings occur. The probability of scattering through an angle greater than φ is simply the ratio of the total target area to the total foil area—that is, $n\pi b^2 tA/A$ or $n\pi b^2 t$. Thus, if we have N_i incident alpha particles, the number scattered through an angle greater than φ, N_s, is given by $N_s = N_i(n\pi b^2 t)$, or the fraction scattered is $N_s/N_i = n\pi b^2 t$.

All factors in the above equation, with the exception of b, can be determined experimentally. With the aid of the calculus, a relationship between b and φ can be derived. When this is done one gets the following:



$$\frac{N_s}{N_i} = \frac{(9 \times 10^9)\ Z^2 q_e^4 ntA}{4R^2K^2\ \sin^4\phi/2}$$

where Z = the atomic number of the scattering nuclei

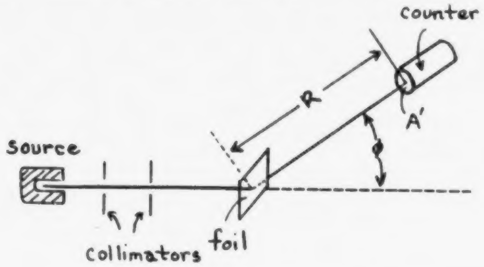
q_e = the electronic charge

A = the area of the counter window

R = the distance from the foil to the counter

K = the kinetic energy of the alpha particle.

With this equation, which is essentially the relation given to Geiger by Rutherford, experimental tests on the nuclear model could be made. It was found that for both gold and silver foils there was agreement between theory and experiment.



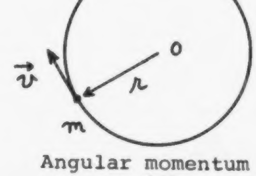
The above scattering formula is seen to be proportional to Q^2 and inversely proportional to the square of the alpha particle's kinetic energy (or v_α^4) and to $\sin^4\phi/2$. It has been stated that Rutherford did not derive the above scattering formula. According to George Gamow,

> Rutherford was so poor in mathematics that the famous Rutherford formula for alpha particle scattering was derived for him by a young mathematician, R. H. Fowler.

For a complete derivation of the Rutherford scattering law see Introduction to Atomic and Nuclear Physics, Henry Semat, Holt, Rinehart, and Winston, Fourth Edition, 1963, Appendix VII, page 598.

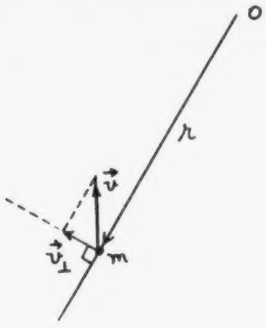
Angular Momentum

The angular momentum of a particle is always defined with respect to a point. Sometimes this point is called the center of rotation. In the case of circular-type motion it is particularly simple; i.e., the point is the center of the circle. In circular motion it is also simple because the velocity, v, is always perpendicular to the radius, r. In this case the angular momentum can be represented as mvr, where m is the mass. The situation is pictured below.

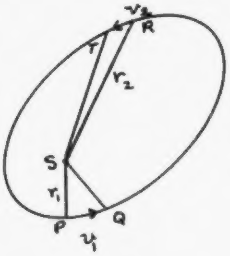


Angular momentum of mass m with respect to center O is mvr.

In the more general case, the angular momentum is written as mv r, where v indicates the component of v that is perpendicular to r, the distance from the point of reference to the object. This more general case is illustrated below.



As a point of interest, angular momentum is conserved for a system just as is linear momentum. In fact, Kepler's second law, the law of equal areas, is an expression of the conservation of angular momentum. This can be shown as follows. Kepler's second law states that for equal transit times between PQ and RT,



$$A_1(\Delta PQS) = A_2(\Delta RTS).$$

When transit time between PQ and RT is small, arc $\widehat{PQ}$ = chord $\overline{PQ}$. Now

$$A_1 = \tfrac{1}{2}r_1(\overline{PQ})_\perp$$

where $(PQ)_\perp$ is the component of $\overline{PQ}$ perpendicular to r_1.

Likewise, $A_2 = \tfrac{1}{2}r_2(\overline{RT})_\perp$. But, $\overline{PQ}\alpha v_1$ and also $\overline{PQ}\alpha mv_1$; thus, $(PQ)_\perp \alpha (mv_1)_\perp$. Substituting, we get $A_1 \alpha \tfrac{1}{2}r_1(mv_1)_\perp$. In like manner, $A_2 \alpha \tfrac{1}{2}r_2(mv_2)_\perp$. Since $A_1 = A_2$, proportionality constants will cancel and

$$mv_{1_\perp}r_1 = mv_{2_\perp}r_2$$

which says that the product, $mv_\perp r$, the angular momentum, is a constant; that is, angular momentum is conserved.

Nagaoka's Theory of the "Saturnian" Atom

In the same volume of the Philosophical Magazine in which this [Thomson's] paper appeared, another theory of atomic structure was propounded by H. Nagaoka. This was the theory of the 'Saturnian' atom and was the precursor of the nucleus theory so brilliantly developed by Rutherford and Bohr. Nagaoka's theory derived its inspiration from the mathematical analysis of the stability of the system of rings surrounding the planet Saturn, by James Clerk Maxwell who, in 1856, was awarded the Adam's Prize for an essay, entitled 'On the Stability of the Motion of Saturn's Rings', ... In this essay, Maxwell discusses the stability first of a solid ring of matter and then of a ring consisting of a number of separate particles; in both cases the rings were presumed to be rotating around, and to be attracted inversely as the square of their distances from, a massive central body. He concluded that although the system containing a solid ring would be unstable, a massive central body surrounded by a ring of separate satellites would form a stable system if the angular velocity of the ring were sufficiently high. ...

Such a stable structure of separate particles rotating in a series of concentric rings round a massive central body which attracts the satellite particles with a force inversely proportional to the square of their radii of rotation was suggested as a possible model of the atom by H. Nagaoka in 1903 in a paper read before the Physico-Mathematical Society of Tokyo. This was published in the Phil. Mag., Vol. 7, p. 445, 1904, and was entitled 'Kinetics of a System of Particles illustrating the Line and Band Spectrum and the Phenomena of Radioactivity'. ...

In order to account for the characteristic frequency lines of the band spectrum Nagaoka supposed that the rings of electrons in the 'Saturnian' atom would vibrate and that these vibrations would give rise to radiation. ...

"There are various problems which will possibly be capable of being attacked on the hypothesis of a Saturnian system, such as chemical affinity and valency, electrolysis and many other subjects connected with atoms and molecules. The rough calculation and rather unpolished exposition of various phenomena above sketched may serve as a hint to a more complete solution of atomic structure."

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*Excerpted from Conn and Turner, The Evolution of the Nuclear Atom, Iliffe Books Ltd., London, 1965, pp. 111-118.

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1963	Sept	Some Observations on the Theory of Electrons and Atomic Nuclei in Solids (W.V. Houston)
1964	July	Low Energy Electron Diffraction (L.H. Germer)
1966	Oct	Atomic Lifetimes and Electron Excitation (H.H. Stroke)
1967	May	The Septuagenarian Electron (G. Thompson) During the last 70 years, concepts of the electron have developed and changed greatly.

American Journal of Physics

1966	Apr	Visual Observation of Low Energy Electron Beams (D. Haneman)
1966	Apr	Are Electrons Real? (W.W. Houston)
1966	Mar	Intensity of Electron Beams (J.C. Helmer)

Physics Teacher

1963	Feb	Electron Diffraction: Discovery (C.J. Calbick)
1965	Jan	Charging an Electroscope - methods
1965	Apr	Charging an Electroscope - methods
1966	Sept	Electrons, Photons and Students (A.M. Portis)
1967	Apr	Magnetic Experiments with a Cathode Ray Tube (J.G. Shepherd)

RELATIVITY

Scientific American

1963	Feb	The Clock Paradox (J. Bronowski) The famous result of relativity is explained in terms of the Pythagorean theorem.

American Journal of Physics

1964	July	Speed and Kinetic Energy of Relativistic Electrons (W. Bertozzi)
1965	July	The Geometrical Appearance of Large Objects Moving at Relativistic Speeds (G.D. Scott and M.R. Viner)
1966	July	The Relativistic Rocket (K.B. Pomeranz)

Physics Teacher

1964	Apr	The Special Theory (R.P. Feynman) From the Feynman Lectures
1965	May	Time Dilatation, Space Contraction (D. Kutliroff) Relativity for high school students.
1966	Jan	Time (G.J. Whitrow) - relativity

SPECTRUM-STRUCTURE

Scientific American

1964	Mar	Fast-Neutron Spectroscopy (L. Cranberg) New techniques make it possible to use energetic neutrons to probe the nucleus.
1965	Mar	The Structure of Crustal Surfaces (L.H. Germer) The arrangement of their atoms as revealed by probing with low-energy electrons.

Physics Today

1965	Aug	Cross-Section Measurements (A. Hemmendinger) Made with neutrons from a nuclear detonation.
1966	Oct	Nuclei of Low to Medium Mass (G.I. Harris and P. Goldhammer) Many structural details are already studied for mass numbers between 16 and 56.
1967	May	Nuclear Structure and Modern Research (V.F. Weisskopf) Our understanding of nuclear physics plays a growing role in science and technology.

American Journal of Physics

1966	May	Descartes on the Refraction and the Velocity of Light (J.G. Burke)
1966	May	Maxwell's Orals and the Refraction of Light (M.H. Sussman)
1966	Sept	Simple Mossbaver Spectrometer Using X-Ray Film (E. Kankeleit)
1966	Nov	X-Ray Diffraction and the Bragg Law (Elton and Jackson)
1966	Nov	Spectrographic Analysis with a Small Telescope and Transmission Grating (Warren and Graedel)
1967	Apr	Interference in Scattered Light (deWitte)

Science (Weekly)

1967	Mar	Ratio of Blue to Red Light: A Brief Increase Following Sunset (T.B. Johnson et al.)

Physics Teacher

1964	Feb	Measuring the Wavelength of Light (M.L. Clark) Slits of known distance apart.
1964	Oct	Spectra Inform Us about Atoms (W.F. Meggers)
1965	Oct	History of X-Ray Analysis (Sir Lawrence Bragg)
1965	Apr	The Rainbow
1966	Nov	Determining Light Wavelengths Individually by Use of One Special Source for the Class (H.H. Gottlieb)
1966	Mar	A Simple Demonstration Spectroscope (Z.V. Harvalik) - four prisms, a projector, and screen
1967	Jan	Mass Spectroscopy - An Old Field in a New World (A.O. Nier)

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1965	Jan	Infrared Spectroscopy (R.N. Jones) About the most powerful molecular probes to tell about structure.

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1965	Jan	Allesandra Volta (Giorgio de Santillana) A review of his work evokes the excitement of the first discoveries in electricity.

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1963	Jan	No Fugitive and Cloistered Virtue (J.A. Wheeler) A tribute to Niels Bohr.
1963	Oct	Niels Henrik David Bohr (1885-1962) Memories of Niels Bohr (J.R. Nielsen) Reminiscences of Niels Bohr (F. Blotch) Remarks at Niels Bohr Memorial Session (A. Bohr) Niels Bohr and Nuclear Physics (J.A. Wheeler) Niels Bohr's Contribution to Epistemology (L. Rosenfeld) Niels Bohr: A Memorial Tribute (V.F. Weisskopf)
1965	Jan	Einstein and Some Civilized Discontents (M.J. Klein) Biography.
1966	Sept	The Two Ernests (M.L. Oliphant) The author recalls Rutherford and Lawrence in the early days of nuclear research.
1966	Oct	The Two Ernests - II (M.L. Oliphant) Rutherford and Lawrence, in lively letters, reveal the growth of nuclear physics.
1967	Apr	Nagaoka to Rutherford, 22 February 1911 (L. Badash) A Japanese physicist describes his "grand tour" of European physics laboratories.

American Journal of Physics

1964	Sept	Rutherford and his Alpha Particles (Osgood and Hirst)
1964	Nov	Millikan - Teacher and Friend (H.V. Neher)
1965	Feb	Anniversaries in 1965 of Interest to Physicists (E.S. Barr) a) Louis Carl Heinrich Friedrich Paschen b) Pieter Zeeman Biographical sketches.
1966	Jan	Anniversaries in 1966 of Interest to Physicists (E.S. Barr) Michael Faraday - biographical sketch.

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1963	Nov	Faraday's Diary
1964	Jan	Robert Andrews Millikan (E.C. Watson) - biography
1964	Nov	Millikan, Teacher and Friend (H.V. Neher)
1965	Feb	Michael Faraday (L.P. Williams) - biographical sketch
1965	Apr	H.G.V. Moseley (L.A. Redman)

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Physics Teacher

1966	Nov	The Key Ideas of Quantum Mechanics (K.W. Ford)

SPECTRUM

Physics Today

1963	Mar	Absorption Spectroscopy (Bauman) A bit rough.
1963	May	An Introduction to Infrared Spectroscopy (Brugel)
1963	Oct	Fifty Years of X-Ray Diffraction (Ewald, ed.)
1963	Nov	Nuclear Spectroscopy (Proc. of the International School of Physics)
1964	Apr	X-Ray Diffraction in Ice Crystals, Imperfect Crystals, Amorphous Bodies (Goinier)
1964	Nov	X-Ray Optics and X-Ray Microanalysis (Pattee et al., eds.)
1965	May	Atomic Spectra and the Vector Model (Candler)
1965	Sept	Spectrum of Thorium from 9400 to 2000 Å (Junkas and Salpeter)
1965	Sept	Spectroscopic Properties of the Rare Earths (Wyborne)
1965	Oct	Alpha-, Beta-, and Gamma-Ray Spectroscopy (Siegbahm, ed.)
1965	Dec	Atomic Spectroscopy in the Vacuum Ultra Violet from 22500 to 1100 Å Part I (Junkes, Salpeter, Milazzu)

Physics Teacher

1965	Mar	Experimental Spectroscopy (R.A. Sawyer)
1965	Feb	Optics, or a Treatise of the Reflections, Refractions, Inflections and Colors of Light (Sir Isaac Newton)
1965	Nov	Optics, Waves, Atoms, and Nuclei: An Introduction (E.L. Goldwasser)
1966	Apr	Introduction to Molecular Spectroscopy (A.J. Sannessa)

HISTORY

Physics Today

1963	Dec	Rutherford at Manchester (Birks, ed.)
1965	Apr	J.J. Thompson and the Cavendish Laboratory in his Day (Thompson's son)
1966	May	The Collected Papers of Lord Rutherford of Nelson (Cavendish Laboratory)
1967	Jan	Robert Boyle on Natural Philosophy (M.B. Hall) About Boyle's various scientific interests.

American Journal of Physics

1966	Feb	Michael Faraday (L.P. Williams)

Physics Teacher

1964	Oct	Faraday, Maxwell, and Kelvin (D.K.C. MacDonald)
1964	Oct	Rutherford and the Nature of the Atom (Andrade)
1966	Feb	Michael Faraday (L.P. Williams) - biography

Suggested Answers to Unit 5 Tests

Test A

ITEM	ANSWER	SECTION OF UNIT
1	D	20.1
2	E	18.4, 20.1, 20.2
3	A	18.2
4	B	general
5	A	18.4
6	A	17.8
7	D	18.6, 20.1
8	A	19.4, 19.5
9	B	17.6
10	C	18.2
11	A	19.2
12	A	18.2
13	C	20.1
14	B	20.5
15	A	20.1

Answers
Test A

Group I

1. Section of Unit: 18.2

The particle has a mass that is 1836 times the mass of an electron. The particle has a charge that is 1/1836 the charge of the electron. The mass and charge of the particle are different from the mass and charge of an electron.

2. Section of Unit: 17.2

$$\begin{aligned} \text{Molecular mass of ZnO} &= \text{Atomic mass of zinc} + \text{Atomic mass of oxygen} \\ &= 65.37 + 15.99 \\ &= 81.36 \\ \text{\% by mass of zinc} &= \frac{(65.37)}{(81.36)}\ (100\%) \\ &= 80.3\% \end{aligned}$$

3. Section of Unit: 18.2

The deflection of the electron beam in a magnetic field depends upon the direction of the beam relative to the magnetic field, the speed of the electrons in the beam, and the strength of the magnetic field.

4. Sections of Unit: 19.4, 19.5

The term "mvr" is the angular momentum of an electron as it orbits about the positive nucleus of a hydrogen atom. Equating the electron's angular momentum with the term "nh"/2π indicates that the angular momentum of the electron must be quantized, for n is an integer, whereas h/2π is constant.

5. Section of Unit: 20.3

$$\text{de Broglie}\quad \lambda = \frac{h}{mv}$$

$$\lambda = \frac{6.6 \times 10^{-34}\ \text{J-sec}}{(1.67 \times 10^{-27}\ \text{kg})\ (10^{8}\ \text{m/sec})}$$

$$\lambda = 3.95 \times 10^{-15}\ \text{m}$$

Group II

6. Sections of Unit: 19.4, 19.5

Balmer

$$1/\lambda = R_H \quad \frac{1}{n_f^{\,2}} - \frac{1}{n_i^{\,2}}$$

Bohr

$$hf = \frac{E_1}{n_f^{\,2}} - \frac{E_1}{n_i^{\,2}}$$

$$h\,\frac{c}{\lambda} = E_1 \left(\frac{1}{n_f^{\,2}} - \frac{1}{n_i^{\,2}}\right)$$

$$\frac{1}{\lambda} = \frac{E_1}{hc} \left(\frac{1}{n_f^{\,2}} - \frac{1}{n_i^{\,2}}\right)$$

$$R_H = \frac{E_1}{hc}$$

$$\frac{1}{\lambda} = R_H \left(\frac{1}{n_f^{\,2}} - \frac{1}{n_i^{\,2}}\right)$$

7. Section of Unit: 20.5

a) The uncertainty principle states that it is not possible to measure simultaneously an electron's position and velocity (momentum) to any prescribed accuracy.

b) $(\Delta x)\ (\Delta p) \geq \frac{h}{2\pi}$

$$\Delta p \geq \frac{h}{2\pi} \left(\frac{1}{\Delta x}\right)$$

$$\Delta p \geq \frac{6.6 \times 10^{-34}\ \text{J-sec}}{2(3.14)\ (10^{-10}\ \text{m})}$$

$$\Delta p \geq 1.05 \times 10^{-24}\ \text{kg m/sec}$$

Suggested Answers to Unit 5 Tests

Test B

ITEM	ANSWER	SECTION OF UNIT
1	D	20.1
2	D	general
3	C	19.1
4	A	18.2
5	B	18.6
6	A	20.4
7	E	general
8	D	19.2
9	E	18.6
10	A	20.6
11	C	20.3
12	B	general
13	C	20.4
14	A	18.3
15	B	17.3

Answers
Test B

Group I

1. Section of Unit: Chapter 20

One can attribute both wave-like behavior and particle-like behavior to everything in the universe. For example, the diffraction of an electron can be explained by its wave properties and its momentum can be explained by its mass and velocity properties.

2. Section of Unit: 19.8

i) Although the Bohr model accounted for the spectra of atoms with a single electron in the outermost shell, serious discrepancies between theory and experiment appeared in the spectra of atoms with two or more electrons in the outermost shell.

ii) Bohr's theory did not account in a quantitative way for the splitting of spectral lines that occurred when the sample being studied was in an electric or magnetic field.

iii) Bohr's theory supplied no method for predicting the relative intensity of spectral lines.

3. Section of Unit: 18.4

The energy of a photon is directly proportional to its frequency. This can be said more succinctly by means of the equation $E = hf$, where h, the constant of proportionality, is Planck's constant.

4. Sections of Unit: 17.5, 17.6

i) The periodic table provided through the introduction of a system of "numerical characterization" of the elements a dependable means of correlating the elements and their properties. It established the regular occurrence of physical and chemical properties, and suggested some periodic recurrence of structure in atoms.

ii) Gaps in the periodic table led Mendeleev to predict the existence of undiscovered elements, and furthermore allowed him to describe accurately many of their properties.

5. Prologue

Alchemy, the futile attempt to transmute base metals into gold, was the forerunner of modern chemistry. Its importance lies in its by-products such as the development of methods of chemical analysis, the study of the properties of many substances and processes such as calcination, distillation, fermentation and sublimation, and the invention of many pieces of chemical apparatus that are still used today.

6. J. J. Thomson's e/m experiment —— 18.2

Thomson showed that cathode-ray particles (electrons) were emitted by many different materials. Their charge was similar in magnitude to that of a hydrogen ion, but they were considerably less massive than the hydrogen ion. He concluded that these particles form a part of all kinds of matter, and in so doing suggested that the atom is not the ultimate limit to the subdivision of matter.

Millikan's oil-drop experiment —— 18.3

Millikan's experiment, by showing that the electric charge picked up by an oil drop is always an integral multiple of a certain smallest value, demonstrated that charge is quantized.

Photoelectric experiments —— 18.4

These experiments show that the maximum kinetic energy of photoelectrons increases linearly with the frequency of the incident light, provided the frequency is above the threshold frequency. This threshold frequency is different for different metals. Photoelectrons are emitted at frequencies just above the threshold no matter how low the intensity of the incident light. In addition, there is practically no time lag between the instant the incident light strikes the target and the emission of photoelectrons. At frequencies just below the threshold no electrons are emitted no matter how intense the incident light. In summary, these experiments show that the energy of light is a function of its frequency and that light energy is quantized.

Faraday's electrolysis experiments —— 17.7, 17.8

Faraday's experiments showed that a given amount of electric charge is closely related to the atomic mass and valence of an element and in so doing implied that, a) matter is electrical in nature, and b) that electricity is atomic (quantized) in nature.

Rutherford's alpha-particle scattering experiments —— 19.2

Rutherford's alpha-particle scattering experiments showed that the atom is mostly empty space. The experiments demonstrated that there was a positive charge within the atom that occupies a very small amount of space, and furthermore it is this charge that scatters alpha particles by a coulomb force of repulsion.

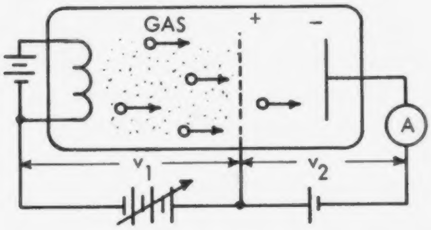
7. Sections of Unit: 17.1, 17.2

   a) Because atoms are unchangeable, Dalton inferred that matter must be conserved, that all substances must be composed of different arrangements of atoms, and that atoms combine in different groups.

   b) The weight of A relative to B is 6 to 1. However, in a compound of only A and B it is found that there is 3 times as much A as there is B (by weight). Consequently, the only possible formula for this compound is AB_2.

8. Section of Unit: 19.6

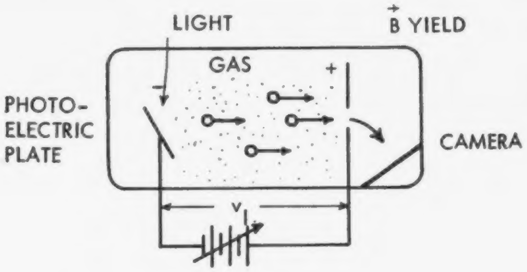
The following is a greatly simplified schematic diagram of the Franck-Hertz apparatus. For the purposes of this question it should be considered a more than adequate answer.



v_1 = accelerating potential

v_2 = small stopping potential

However, devices other than those shown in the diagram above can be used to make the necessary measurements. For example, the following diagram shows a satisfactory arrangement.



v = accelerating potential

The amount of bending of the electron beam is inversely related to the kinetic energy of the emerging electrons.

A workable apparatus would have to include:

1. a source of electrons
2. a means of accelerating the electrons
3. a chamber where the electrons pass through a gas
4. an electron detector (ammeter)
5. some means of measuring the kinetic energy of the electrons after they emerge from the gas

Suggested Answers to Unit 5 Tests

Test C

ITEM	ANSWER	SECTION OF UNIT	PROPORTION OF TEST SAMPLE ANSWERING ITEM CORRECTLY	ITEM	ANSWER	SECTION OF UNIT	PROPORTION OF TEST SAMPLE ANSWERING ITEM CORRECTLY
1	A	19.2	.46	21	E	17.1, 17.2	.76
2	C	20.1	.42	22	A	19.1	.52
3	A	19.2	.65	23	A	18.3	.72
4	E	20.2	.32	24	B	17.7, 17.8	.64
5	D	20.3	.61	25	D	17.7, 17.8	.67
6	B	19.4, 19.5	.53	26	D	17.3	.58
7	A	17.7, 17.8	.66	27	C	Chapter 20	.74
8	C	19.4, 19.5		28	D	18.4	.53
9	C	18.4	.70	29	C	18.6	.42
10	A	18.4	.45	30	B	general	
11	B	19.2	.67	31	E	20.1	.48
12	B	general		32	C	20.4	.45
13	B	19.4, 19.5	.80	33	B	19.4	.54
14	C	19.6	.38	34	A	Chapter 20	.71
15	E	18.2	.53	35	E	19.2	.62
16	C	19.4		36	C	17.7, 17.8	
17	B	19.2	.71	37	D	20.1	.76
18	A	18.4		38	B	20.5	.78
19	B	18.6	.48	39	A	19.4	.71
20	A	18.2	.65	40	A	19.6	.59

Suggested Answers to Unit 5 Tests

Test D

1. Section of Unit: 17.8

This argument has three steps. The question provides the first and the last——the student must apply the missing middle step.

STEP 1. A given amount of electric charge is related to the atomic mass and valence of an element.

STEP 2. Atomic mass and valence are characteristics of the atom of the element.

STEP 3. Therefore, a certain amount of electric charge is associated with an atom of the element.

2. Section of Unit: 18.3

Millikan's experiment, in showing that the electric charge picked up by an oil drop is always an integral multiple of a certain minimum value, demonstrated that charge is quantized.

3. Section of Unit: 18.5

Rutherford could not explain the bright-line spectrum of hydrogen. In addition, he had nothing to say about the details of distribution of negative charge. Bohr explained the bright-line spectrum of hydrogen by suggesting that a spectral line may be attributed to the quantized release of energy by the change in the nucleus-electron energy state. Bohr also described in some detail the orbiting electron in terms of permissible quantized distances from the center of the nucleus, quantized angular momentum and quantized energy.

4. Section of Unit: 18.5

Einstein's formula: $KE_{max} = hf-W$

When the emitted photoelectron has practically no KE, incident light with a minimum f has caused this emission. Therefore,

$$hf = W$$

$$f = \frac{W}{h}$$

$$f = \frac{2 \times 10^{-18}\ J}{6.6 \times 10^{-34}\ J\text{-sec}}$$

$$f = 3 \times 10^{15}\ sec^{-1}$$

5. Section of Unit: 20.1

As the speed of an electron increases, its mass increases without limit. This can be stated mathematically as follows:

$$m = \frac{m_o}{\sqrt{1 - \frac{v^2}{c^2}}}$$

6. Section of Unit: 17.8

Faraday's second law of electrolysis states that 96,500 coulombs will produce 1.00 grams of hydrogen. The problem indicates that a current of 3 amperes flows through water for 60 minutes (3600 sec.). This is equivalent to the passage of 9800 coulombs ($q = I \cdot t$)

The mass of hydrogen produced = $(1.00)\ \frac{(10,800)}{(96,500)} = 0.112$ g.

The ratio of the amounts of oxygen and hydrogen liberated in the electrolysis of water is 8:1.

Therefore, the mass of oxygen produced = (8) (0.112) = 0.896 g.

7. $KE_{max} = hf\text{-}W$ —— 18.5

Einstein showed that photoelectric emission could be explained by the quantization of light energy, thus paving the way to quantum mechanics.

$$\frac{1}{\lambda} = R_H \left(\frac{1}{{n_f}^2} - \frac{1}{{n_i}^2}\right) \quad \text{—— 19.1}$$

Balmer's empirically derived formula summarized some regularity and predicted the existence of other spectral line series in the hydrogen spectrum. Agreement with the predictions of this formula led to Bohr's theoretically derived model of the atom.

$$m = \frac{m_o}{\sqrt{1 - \frac{v^2}{c^2}}} \quad \text{—— 20.1}$$

This equation, showing the relationship between an object's mass and its speed, is one of the more popular consequences of Einstein's special theory of relativity. The introduction of this theory prompted a total reassessment of most areas of physics involving the study of objects in motion. It has been immensely important in the field of high-energy physics, where very small particles are accelerated to speeds that approach the speed of light, and also in astronomy and atomic and nuclear physics.

$$\lambda = \frac{h}{mv} \text{ --- } 20.3$$

The de Broglie equation suggested that matter has wave properties. This prediction found experimental verification when Davisson and Germer demonstrated that electrons could be diffracted.

$$(\Delta x)\ (\Delta p) \geq \frac{h}{2\pi} \text{ --- } 20.5$$

The Heisenberg uncertainty principle states that we are unable to measure simultaneously the position and velocity of an electron. This same reasoning holds for all moving objects, but is of no practical consequence for relatively massive objects. Heisenberg's principle, one of the early consequences of quantum theory, emphasized the nondeterministic probabilistic nature of physics.

8. Section of Unit:

The idea that matter is composed of atoms was proposed by Greek philosophers between 500 and 400 B.C. This theory was devised in an "arm chair" fashion with no attempt at confirming the theory in terms of physical experimentation. The nineteenth century scientists developed theories that they felt would account for the observed properties of matter and tested these theories experimentally. In addition, these atomic theories included a predictive function that would include behavior not yet observed.

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